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CHEMISTRY

FOR IOWA SCHOOLS

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CHEMISTRY FOR IOWA SCHOOLS

Iowa Cooperative Curriculum Development Program

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FOREWORD

This publication for Iowa high schools is part of a science program for grades K-12 developed by committees composed of members representing all educational levels.

More than ever before, future citizens of our country need a broad knowledge of science. This science foundation must be received in the elementary grades and extended throughout the secondary-school level. There is a growing tendency, for those who are college bound, to depend more heavily on the high school chemistry background.

Many new approaches have been developed for the teaching of chemistry. The purpose of this chemistry course guide is to assist the teacher in developing his OWN course so as to best meet the needs of those planning further study, as well as terminal students. It places chemistry in the wider context of science education, and suggests resource material with which to enrich day-to-day teaching.

PAUL F. JOHNSTON
*State Superintendent
of Public Instruction*

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INTRODUCTION

It must be recognized that any well-trained and experienced instructor is capable of developing a satisfactory outline for his own course, given adequate time, library resources, and opportunities for studying curricular trends and the science course organization in his own school system. Indeed it would be presumptuous on the part of any committee to prepare curriculum suggestions were this generally the case. Unfortunately, many instructors lack adequate time for studying the placement of their own course in the wider context of science education as it is currently developing. They may even find time short for gathering the best resources with which to enrich their day-to-day teaching.

This chemistry course guide is an opportunity for the individual instructor to share his burden of course preparation with others in a similar situation. It is his chance to widen his viewpoint through the special studies of others and add to his own experience, the experience of a group. It is on this basis that the following material has been prepared.

The arrangement of topics is only one of several that might be equally satisfactory, yet the initial emphasis on the make-up and structure of chemical substances is believed to be especially sound. The relation of structure to bonding follows logically. The emphasis on chemical bonding as fundamental to a clear understanding of chemical reactions acts as a unifying influence throughout the suggested course. Finally, the topics emphasized are those for which the pupil has been prepared by previous courses in the integrated curriculum of which this guide is a part. Undue overlapping of material already presented, or to be presented later, is avoided while fundamentals which require repetition are given additional emphasis and more mature treatment. Two areas, kinetic theory and electrochemistry, not adequately presented by most high school texts, are given as complete presentations rather than in outline form. Many suggestions for short, vivid demonstrations are made. In areas where demonstrations are not feasible, filmed audio-visual materials are suggested.

The selection of laboratory experiments for a course naturally depends almost entirely on the physical resources of the individual school system and the laboratory time available. Where suggestions for laboratory experiments are made, they are mostly in the form of projects where the think-

ing and planning emphasis is placed upon the pupil. This is in line with the most modern thinking with regard to laboratory exercise.

Since no teaching effort is complete without an assessment of progress toward recognized goals, a short section on evaluation is included.

GENERAL STATEMENT OF AIMS— THE IOWA PLAN

We are in a period in which many new approaches to the teaching of chemistry are being developed for both high schools and colleges. This is good, for a rapidly developing field of science demands new educational methods. Some of these approaches are centered about a core idea as is the Chemical Bond Approach Study mentioned below. Most are tied to a specific book or to a set of class and laboratory texts. This is at once a strength and a weakness of these programs. Many instructors have texts which, on the basis of experience, they believe are superior. In some school systems texts are adopted for a fixed period of time and year-to-year changes are not feasible. The curriculum suggestions made in this outline are not tied to any single text but may be followed with any one of several popular secondary school texts recently published.

In developing the Iowa Plan, so called for want of a more descriptive title, the committee was influenced by several general principles.

Future citizens of our country will have even greater need for a broad viewpoint on scientific matters than those now directing its destinies. Hence, a general coverage of the field of chemistry and soundness of approach to this are both necessary.

The trend in freshman college teaching is to depend more heavily on a high school chemistry background than in the past. A better science foundation received in the elementary grades and greatly improved teaching of chemistry on the secondary-school level makes this possible. At the same time this trend increases the responsibility of the high school for giving its college-oriented pupils an especially well-integrated, basic chemistry course.

Any suggestions on chemistry curriculum, while serving as a guide to the new instructor, should allow the experienced instructor full opportunity to continue the detailed methods of presentation which he has found most valuable in the past.

The Iowa Plan endeavors to implement these

principles by outlining a somewhat traditional course in general chemistry, but with special attention to the early presentation of modern theory and to the properties and reactions of the more important inorganic substances. The emphasis is on understanding rather than on mere acquisition of information and, in the laboratory, upon the development of chemical thinking and precise observation in addition to methodology.

POINTS FOR REVIEW

Review the background material based on junior high physical science and similar units. This should not take more than the first week of school; however, the time allotted should be determined on the basis of a pretest, such as the physical science aptitude test available from the State University of Iowa or other similar tests at the discretion of the instructor. Its purpose is to bring the class together and to recall some of the basic material upon which the chemistry unit is built.

- A. The nature of the universe.
- B. The metric system.
- C. The concepts of mixtures and compounds.

Note: The difference between mixing and combining can be shown by putting a mixture of two parts hydrogen or natural gas with one part of oxygen into a syrup can equipped with a spark plug placed in the side of the can. Do not cover tightly. Use an induction coil to furnish the spark which will cause combustion and combine the gases. *Care! Stand at a safe distance from explosion.*

- 1. Mixtures and their properties.
- 2. Compounds and their characteristics—how they differ from mixtures—the nature of chemical change—the law of definite proportions.
- D. Matter and energy—selected review.
 - 1. Forms of energy—relation to work—principle of interconvertibility—dimensions in which expressed.
 - 2. Sources of energy.
 - 3. Transformations of energy in relation to chemical change.

ARRANGEMENT OF MATERIAL

The following brief outline sets forth the sequence of suggested coverage that is contained within this handbook. The first part of the outline includes the introduction as well as the recall of certain fundamentals basic to the chemistry course.

- I. The behavior of matter
 - A. Matter and energy
 - B. Kinetic theory
 - C. The gaseous state
 - D. The liquid state
 - E. The solid state
- II. Atoms and molecules (an expanded unit)
 - A. Atomic structure
 - B. Chemical bonding of molecules
- III. Chemistry of common gases
 - A. Oxygen and its compounds
 - B. Hydrogen and its compounds
- IV. Formulas and equations
 - A. Meaning and use of formulas
 - B. Naming of inorganic compounds
 - C. Equations
 - D. Problems based upon equations
- V. Water and solutions
 - A. Physical properties of water
 - B. Chemical properties of water
 - C. Uses of water
 - D. Conservation of water resources
 - E. Solutions, with emphasis on aqueous solutions
- VI. Ionization and ionic solutions
 - A. Acids, alkalis, and salts
 - B. Physical properties
- VII. Oxidation and reduction (an expanded unit)
 - A. Oxidation and reduction
 - B. Oxidation numbers
 - C. Balancing oxidation-reduction equations
- VIII. The periodic table
 - A. The basic pattern
 - B. Divisions of the table—periods, families
- IX. Some common elements and their compounds
 - A. Halogens
 - B. Sulfur and its family
 - C. Carbon—bonding peculiarities and inorganic chemistry
 - D. Nitrogen and its family
 - E. The active metals—alkalis and alkaline earths
 - F. Some light metals—Mg, Al, Zn
 - G. Some heavy metals—Fe, Cu
- X. Equilibrium
 - A. Principles governing equilibrium
 - B. Application of principles—qualitative and quantitative
- XI. Electrochemistry (an expanded unit)
 - A. Production of electricity
 - B. Electric cells—general background
 - C. Electric cells—voltaic and electrolytic
 - D. Theory and design

- E. Electrode potentials
- F. Electrolysis
- XII. Organic chemistry of carbon
 - A. Bonding and structure
 - B. Classes of organic compounds—nomenclature
 - C. Some applications of organic chemistry
- XIII. Nucleonics
 - A. The physical background of nuclear changes
 - B. Nuclear chemistry
 - C. Economic and social impacts

Demonstrations and experiments have been drawn from many sources. Nearly every experiment with glassware and chemicals involves some hazard. In most cases, danger has been recognized by specific directions or cautions for the handling of equipment and chemicals. In choosing experiments for pupil use, the *safety of the pupil* should be the *first* consideration.

The instructor is referred to:

Manufacturing Chemist's Association, Inc. *Guide for Safety in the Chemical Laboratory*. New York: D. Van Nostrand Company, Inc., 1954.

R. D. Macomber, "Chemistry Accidents in High School," *Journal of Chemical Education*, XXXVIII (July, 1961), 367-368.

OTHER PROGRAMS AVAILABLE

The Chemical Bond Approach

The Chemical Bond Approach (CBA) is one of two major studies financed by the National Science Foundation. It is being directed by Laurence E. Strong of Earlham College and Kent Wilson of Tufts University. The activities are presently centered on the Earlham College campus although the study originated in a summer institute at Reed College in 1957.

The concern of the group has been to develop a new approach to the first course in chemistry which would cut across the traditional boundary lines of the subject and at the same time avoid duplication between the high school course and the first course in college. The central theme is that of chemical bonding which has become well enough understood to allow an explanation of most chemical properties in these terms. A wealth of new and exciting subject matter is presented in a way to stimulate thinking and use of the imagination. At the same time, this has necessitated the omission of much of the inorganic chemistry of specific elements of importance. There is also the impression that in seeking exciting material

the group has drawn upon areas of physical chemistry such as thermodynamics and kinetics not usually covered thoroughly until the later years of college.

Both text and laboratory experiments are reproduced in 8½ x 11 inch format. Reprinted readings from the *Journal of Chemical Education*, instructor's aids, examinations, and problems are available.

For further information read: Strong and Wilson, *Journal of Chemical Education*, XXXV (January, 1958), 56, and A. B. Garrett, "New Chemistry," *The Science Teacher*, XXVIII (April, 1961), 15-16. Materials may be obtained from Dr. Laurence E. Strong, Department of Chemistry, Earlham College, Richmond, Indiana.

The Chemical Education

Material Study (CHEM study)

The CHEM study, also under N.S.F. support, has as chairman, Glenn T. Seaborg of the University of California, and is being directed by Arthur Campbell of Harvey Mudd College. It had its origin in a conference of chemistry teachers brought together by the American Chemical Society in 1958.

The goal of the CHEM study group was to "identify the irreducible minimum of basic fundamentals that could and should be taught in the high school course and on which the college course would then be built."¹ The result of this effort appears to be a conservative and well-balanced selection of topics which are presented in the first edition of the text and laboratory manual in a strictly modern manner. Laboratory problems, demonstration problems, reading materials, visual aids, and teacher's guides are, or will ultimately become, available. For further information read: A. B. Garrett, "The New Chemistry," *loc. cit.* Glenn T. Seaborg, "New Currents in Chemical Education," *Chemical and Engineering News*, XXXVIII (October, 1960), 97. Materials may be obtained from Arthur Campbell, Harvey Mudd College, Claremont, California.

The EBF Complete Course in Introductory Chemistry

Note: This is not offered as a substitute curriculum but to inform instructors of this material which is available.

This is a complete course in chemistry on 16 mm sound film prepared under a grant from the Fund

¹ A. B. Garrett, "The New Chemistry," *The Science Teacher*, XXVIII (April, 1961).

for the Advancement of Education under the supervision of an advisory committee of the American Chemical Society. The course is taught by John F. Baxter, Ph.D., Professor of Chemistry at the University of Florida. The course consists of 160 half-hour lecture demonstrations on film with supplementary reading materials in the sense that they may be scheduled either at the beginning or at the end of the course. Twelve films are "special content" films which may be omitted from shorter courses without disruption of the sequence. The material is interesting, and the presentation is clear and forceful. For further information write to Encyclopedia Britannica Films, Inc., 1150 Wilmette Avenue, Wilmette, Illinois.

Individual School Programs

Although the sequence of the Iowa Plan has been carefully arranged, experienced instructors may wish to modify it to fit local situations. This is entirely feasible, but radical changes of content should be carefully considered before being adopted. This is true not only in terms of their absolute importance, but also relative to the overall science program of the school.

In any case, the early introduction of atomic and molecular theory, the nature of chemical bonding as it affects the properties of molecules, and the relation of energy to chemical changes are aspects of the Iowa Plan which should be retained. It is important that these receive proper emphasis in any modern chemistry course.

OUTLINE OF COURSE CONTENT, EXPERIMENTS AND CLASS DEMONSTRATIONS

UNIT I THE BEHAVIOR OF MATTER

Experiments and class demonstrations are inserted where applicable within the outline of course content. This material is separated from the outline by broken lines and is specifically related to the preceding content item.

A. Matter and energy—build more mature concepts on the foundation recalled under Item D of Points for Review (page 2).

1. The nature of energy—energy expressed as a product of a potential factor and a capacity factor.

— — — — —
Conservation of energy—support a heavy pendulum from ceiling about three feet from the wall. Draw the pendulum back against your nose with your head against the wall. Let the pendulum go.

It cannot rise to a greater height than the height from which it started. You are safe if you do not move.

- — — — —
2. Potential energy—weight times height, $mg \times h$.
 3. Kinetic energy—force times distance, $\frac{1}{2}mv^2$.

— — — — —
Kinetic Energy to Potential Energy and Reverse. Toss a ball straight up so that it returns to the hand. Explain that the kinetic energy received from the hand is progressively changed into potential energy as the ball rises. At the peak of its flight all of its energy is potential (energy of position) so that it stands still for a moment before falling. In falling, its potential energy is transferred back into kinetic energy (energy of motion). (See supplementary lecture, "Behavior of Gas Molecules.")

— — — — —
B. Kinetic theory.

1. Postulates.
 - a. Matter is composed of particles called molecules.
 - b. The molecules are continually in motion at any temperature above absolute zero.
 - c. The temperature of a body is a measure of the degree of molecular motion.

The size of a molecule may be determined as follows: Prepare a 0.5 per cent solution of oleic acid dissolved in denatured alcohol. Experimentally determine the number of drops in a milliliter of the oleic acid solution. Pour water into a large shallow tray. Dust the surface with lycopodium powder or talc. Add one drop of the oleic acid solution to the prepared surface. Measure the diameter of the clear space and calculate its area. Knowing the volume of the drop and the concentration of the solution, the number of oleic acid molecules added can be calculated. Assuming a monomolecular layer with the molecules oriented perpendicular to the surface with their acid ends in the water, it is possible to calculate the area covered per molecule. This gives the size of the hydrocarbon chain. Stearic acid or other fatty acids may be substituted for oleic acid. Acids with a double bond in the middle of the chain bend over occupying approximately twice the area of saturated acids. (*Scientific American*, September, 1961.)

- — — — —
2. Consequences of molecular motion.

- a. Brownian movement—a characteristic movement of visible particles suspended in a medium caused by the thermal movement of the molecules.

— — — — —
Brownian movement can be demonstrated in the laboratory if a suitable cell is obtained or constructed. See FIGURE 1.

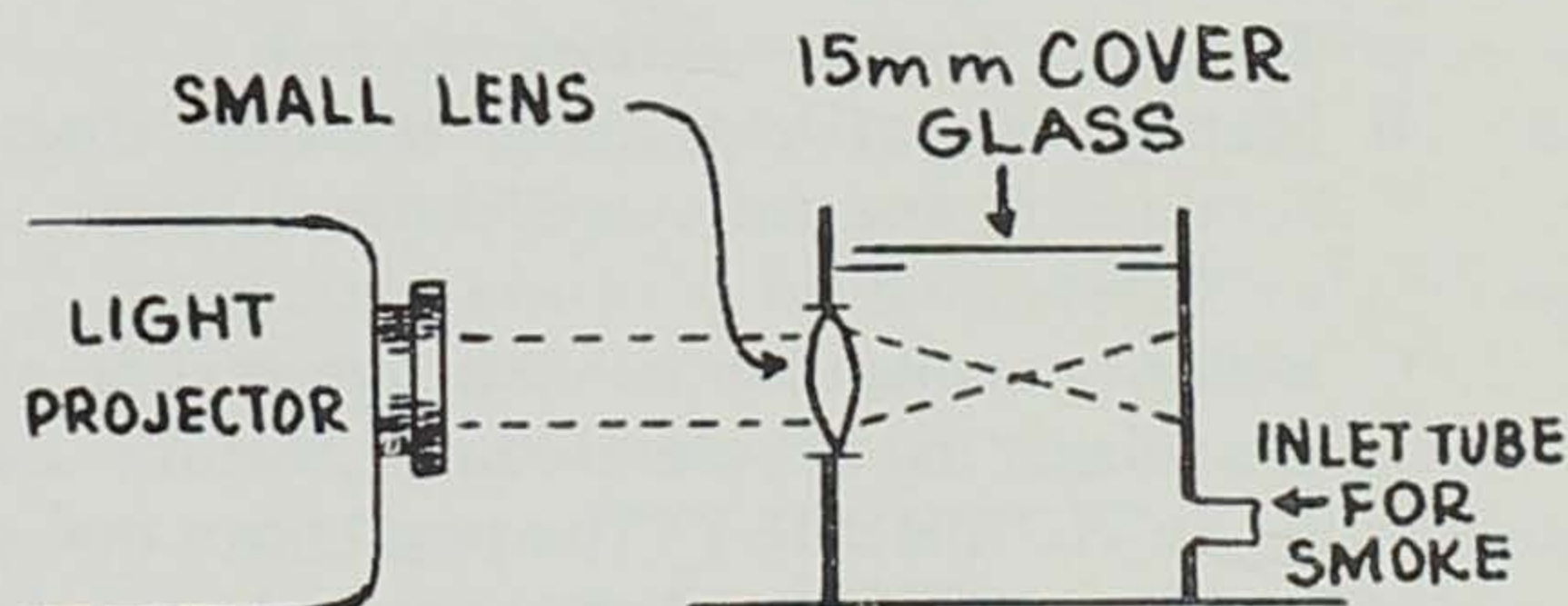


Figure 1

The main requisite is a brightly illuminated closed space into which cigarette smoke can be introduced and viewed by means

of a medium power compound microscope. The cell must be tight enough so that the smoke does not drift across the field of the microscope. Set up the cell as shown in FIGURE 1 and place on the stage of the microscope. Focus the microscope through the cover glass on the point where the rays of light from the projector come to a focus.

- b. Tyndall effect—scattering of light due to the presence of particles in a medium.

Show Tyndall effect by darkening classroom and blowing chalk dust into the beam of a slide projector. Point out that it is the light reflected from the minute particles that is seen, but that the particles themselves are too small to show shape or form.

Tyndall effect may be shown by using a large rectangular glass dish filled with a solution of hypo. A projector should be arranged to send a beam of light through the solution and form a spot on the far wall. Sulfuric acid in very dilute solution is added to the hypo, and the spot on the wall turns blue. As the reaction continues and more particles of elemental sulfur are formed, the color of the spot on the wall changes from blue through yellow and orange to red. This behavior of the fine suspension should be compared with the behavior of a true solution (example: NaCl solution, which will not disperse light—better filter it to make sure).¹

Film: *Molecular Theory of Matter*. I.S.U. NS-2406, 10 min., b & w.

C. The perfect or ideal gas.

1. Its characteristics.

- The basic particle is the molecule.
- Motion of particles is in straight lines, haphazard, and follows Newton's laws.

Make a molecular motion apparatus as described in supplementary lecture on KINETIC THEORY. The pupils can make this as a project or if already made can do experiments with it observing diffusion effects, etc.

¹ Hubert N. Alyea, "Tested Demonstrations in General Chemistry," *Journal of Chemical Education*, XIX (1942), 8, Group VI, The Sulfur Family.

- Distance between molecules is large compared with their size.
 - Molecular motion is proportional to heat energy.
 - Actual gases approach ideal behavior only within certain temperature and pressure ranges.
2. Variables affecting the condition of gases.

Demonstrate properties of gases with equipment available. A 500 ml round bottom flask connected with a mercury manometer can be used to demonstrate Charles' law. Heat the gas in the flask with a burner and show the increase in pressure. Boyle's law can also be demonstrated by compressing the gas in the above flask (at constant temperature) by adding mercury to the flask by means of a dropping funnel in such a way that no gas is lost. The increased pressure which results from the compression shows on the manometer. See FIGURES 2 or 3.

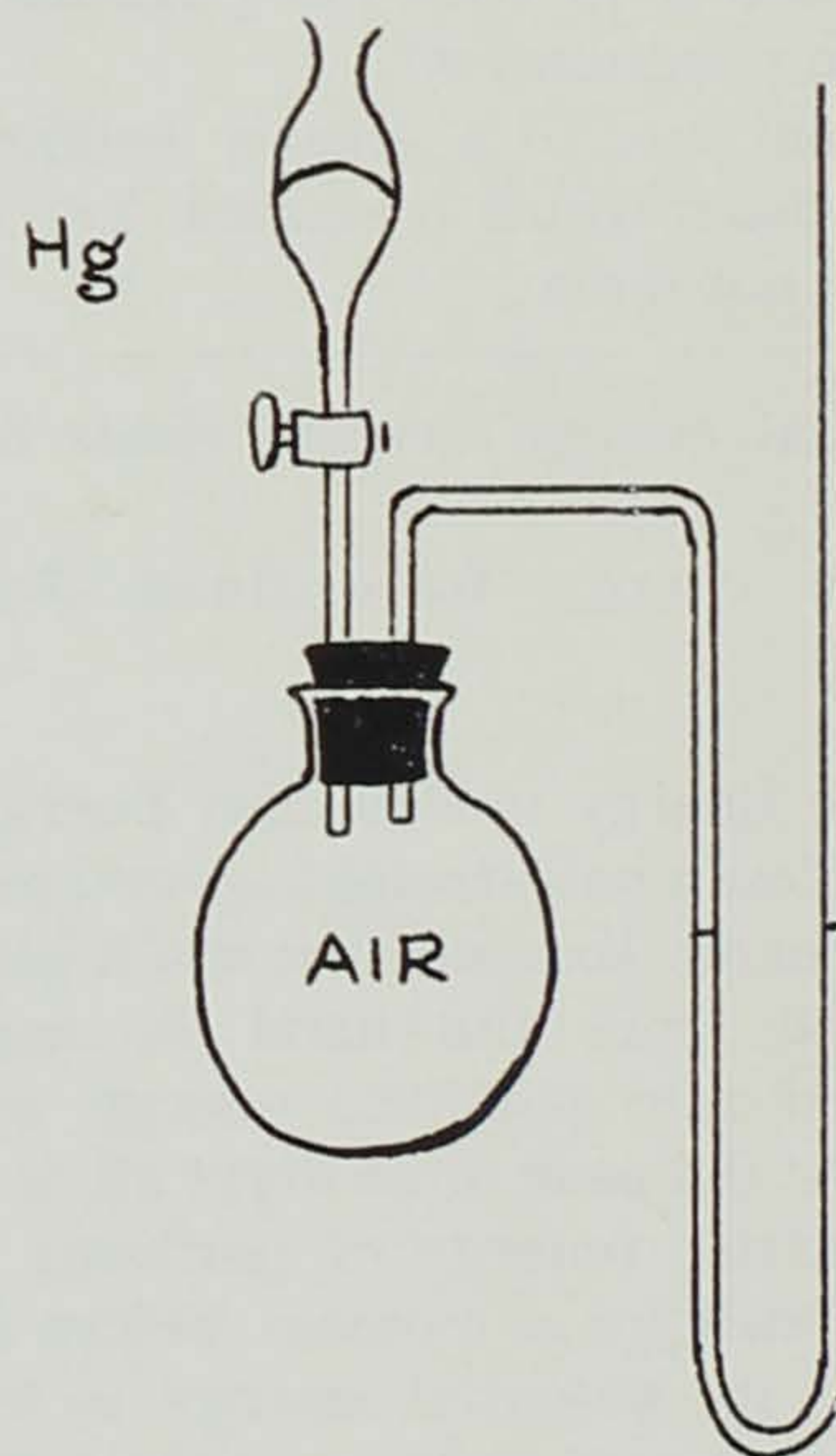


Figure 2

An alternate arrangement for demonstrating Boyle's law is shown in FIGURE 3 which uses a 250 ml graduated cylinder. This has the advantage that the volume of the gas before and after compression can be read from the graduations on the cylinder. It is not satisfactory for the demonstrations of Charles' law since most graduated cylinders will crack if heated.

Pupils can use either setup FIGURES 2 or 3 in the laboratory if a small flask is used, e.g.,

250 or 300 ml r.b. boiling flask. The volume of air used and the volume or weight of mercury used to displace it should be measured. Pupils should calculate from p_1 and v_1 and v_2 after the compression, the value of p_2 and compare with manometer reading.

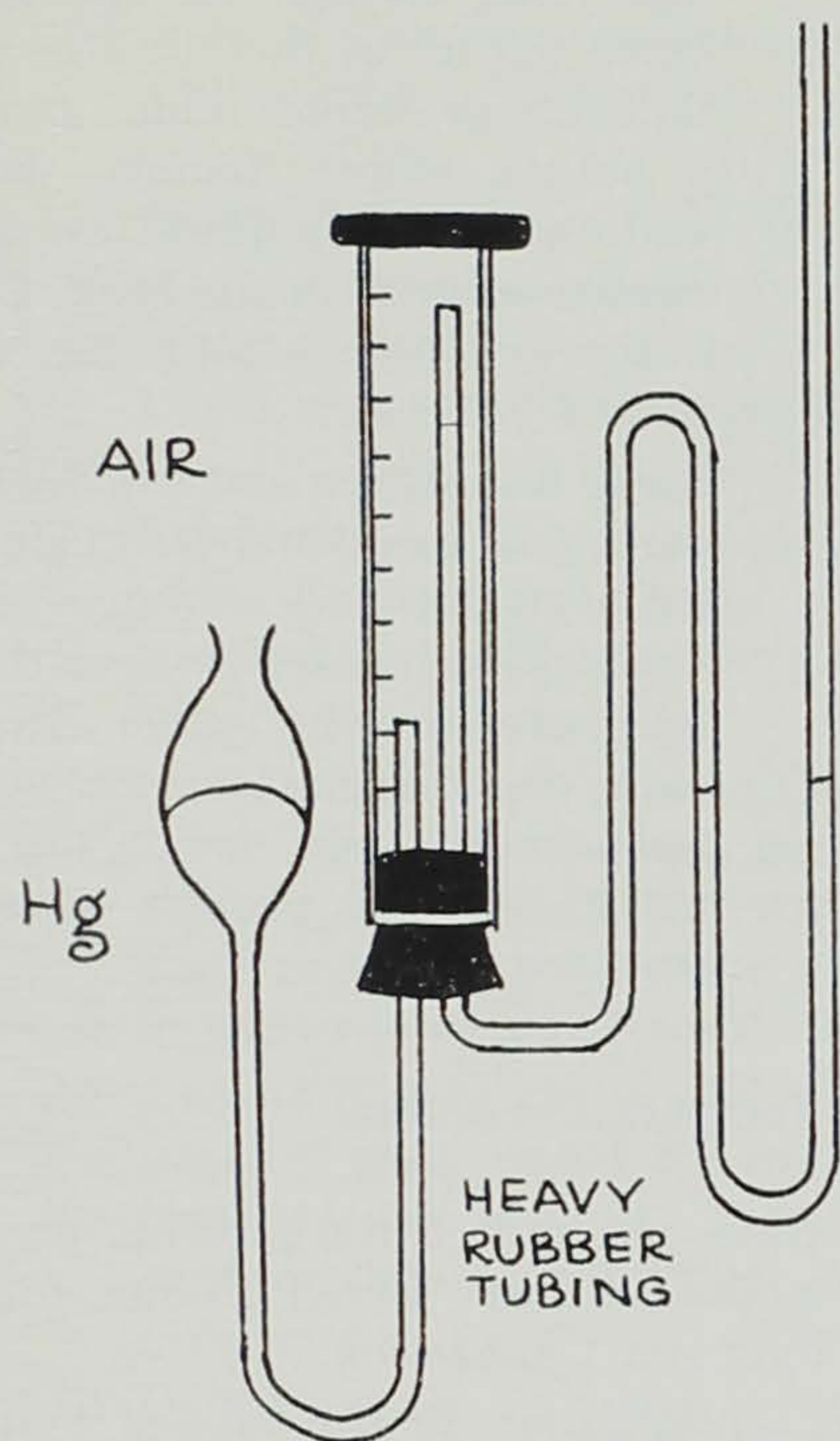


Figure 3

a. Temperature.

Fill a balloon with a gas. Subject it to temperature changes by placing it on ice or very cold water. Then place it over a source of heat. Recall the relation between available energy and molecular motion.

Place 5 ml of water in an erlenmeyer flask. Cover the mouth of the flask with a balloon. Heat the water. Have pupils suggest reason for the change.

Using a J-tube and mercury, have pupils study the relation between pressure and volume of a gas.

Attach a long piece of small bore tubing to an erlenmeyer flask by means of a tight fitting rubber stopper. Place a drop of water or better 50 per cent solution of glycerine, in the tube near the mouth of the

flask so the tube is sealed. Have pupils measure the length of the tube below the drop and call it v_1 . Heat the flask by placing it in a water bath at an elevated temperature measured by a thermometer. When the drop of liquid has moved up the tube and become constant, the distance below the drop should be measured again and called v_2 . Is the change in volume proportional to the absolute temperature change?

- b. Pressure.
 - c. Volume.
 - d. Quantity of gas present.
3. Using the gas theory.
- a. Properties of gases.

Light gases diffuse downward. Fill a wide-mouth bottle with hydrogen. Invert over a like bottle filled with air. While waiting for the diffusion to take place, be democratic and allow the class to vote for what they think is most likely to happen. Possible choices: (1) the hydrogen being lighter will remain in the top bottle, (2) the hydrogen will go to the bottom bottle, (3) the gases will mix together, (4) most of the hydrogen will remain in the top bottle, and (5) do not know. Decide which answer is right by testing each bottle with a candle flame.

Submerge a beaker in a glass container and fill it with water. Invert it. Submerge another beaker so that air is trapped in it. Pour it into the other beaker. Note the fluid nature of the gas.

Fill a small plastic refrigerator bag with Freon and seal. Weigh the gas-filled bag on an analytical balance. Empty the bag and refill with hydrogen or with carbon dioxide and reweigh. Suggest differences in the gases which can account for the observations.

Gases are held in compressed volume by molecular forces or absorption. Fill a large-mouth bottle with granular activated charcoal. Fit the bottle with a two-hole stopper carrying a thistle tube and delivery tube. Announce that this filled bottle contains more than its volume. Proceed to demonstrate by slowly filling the bottle with water while collecting the expelled gases over water in a large graduated cyl-

inder filled with water and inverted over the delivery tube.

-
- (1) Gases can be compressed because there is space between the molecules.
 - (2) Gases can be expanded easily because the forces between molecules are small.
 - (3) Gases diffuse because their molecules are in motion and there is a great deal of space between molecules.
 - (4) The gases as we know them have different densities because the molecules of different gases have different masses.

Standard conditions of T. & P. necessary for comparison of gas densities.

Effect of gas density on sound transmission and other properties of gases can be demonstrated by filling several balloons with different gases such as air, carbon dioxide, natural gas, etc. Keep at same pressure. Fix a whistle to a short piece of glass tubing and attach to each balloon in turn. Note the pitch in each case.

-
- (5) We can change gases into liquids because we can change the distances between molecules by increasing pressure or decreasing the energy of movement.

Have pupils use a distillation set up of flask and water condenser to distill suspensions and solutions. In each case the distillate should be collected in a clean beaker and a portion of this evaporated to dryness so as to observe any solid material distilled with the water. Have the pupils suggest reasons for the results observed. Call attention to the energy relations involved in the process of distillation.

-
- (6) Gases exert pressure because the molecules collide with the surface of the confining container.
 - (7) Vapor pressure is the result of movement of molecules. Those with higher

kinetic energies escape from the liquid into the gaseous state.

Demonstrate vapor pressure phenomena and boiling under reduced pressure. Boil water in a round bottom flask long enough for the steam to drive out all of the air. Use a porcelain chip or carborundum granule for a boiling stone. Remove the flame and immediately place an air-tight rubber stopper in the flask. Cool and observe. Give reasons for the observed behavior.

Some interesting and worthwhile experiments can be devised using a water aspirator and mercury manometer. For example, how does the minimum pressure obtainable with a common water aspirator vary with the temperature of the water? This makes a good project of a semi-research nature.

b. Temperature scales.

Have pupils find fixed points on a thermometer. See "Fixed Points for Thermometer Calibration."¹

Call attention to the factors that might account for any errors observed in the instrument. Emphasize the importance of verifying the calibration of instruments used in exact work. This can be extended to various thermometers and to other similar simple laboratory apparatus.

-
- (1) Common scales.
 - (a) Fahrenheit.
 - (b) Centigrade.
 - (c) Absolute or Kelvin.
 - (2) Conversion between scales.

c. Gas laws.

The gas laws may be illustrated by a device prepared by drilling holes near each end and in the center of a small board about the size of a yardstick. Label each hole appropriately, pressure, temperature, and volume. Show by pivoting at the cen-

¹ *Handbook of Chemistry and Physics* (forty-second edition; Cleveland, Ohio: Chemical Rubber Publishing Company, 1960), p. 2322.

ter hole (temperature) if temperature remains constant, the pressure rises as the volume lowers. Other controls will illustrate other relations.

-
- (1) Boyle's law states that for a fixed mass of gas, the product of pressure and volume is a constant at constant temperature.
 - (2) Charles' law states that for a fixed mass of gas at constant pressure, the volume of a gas is directly proportional to the Kelvin temperature.
 - (3) Gay-Lussac's law states that for a fixed mass of gas at a constant volume, the pressure of a gas is directly proportional to the temperature Kelvin.
 - (4) Gay-Lussac's law of combining volumes states that when gases react their volumes and the volumes of their products, if they are gases, are in the ratio of small whole numbers.
 - (5) Avogadro's law states that equal volumes of gases under the same conditions of temperature and pressure contain the same number of molecules.
 - (6) Dalton's law of partial pressures states that the total pressure of a mixture of two or more gases is equal to the sum of the pressures of the individual gases.
 - (7) Graham's law of diffusion states that the rate of diffusion of a gas is inversely proportional to the square root of its density (or its molecular weight).

Demonstration of effusion through an orifice. Relative rates of effusion follow Graham's law of diffusion. Prepare the modification of Schilling's Apparatus shown in the FIGURE 4. A tall cylinder is needed for the outside shell.

A 40 cm length of 3 cm glass tubing closed with a two-hole rubber stopper makes a satisfactory inner tube. The effusion opening may be a short capillary of glass formed by drawing out a piece of 6 mm tube, but this will give some error due to viscosity.

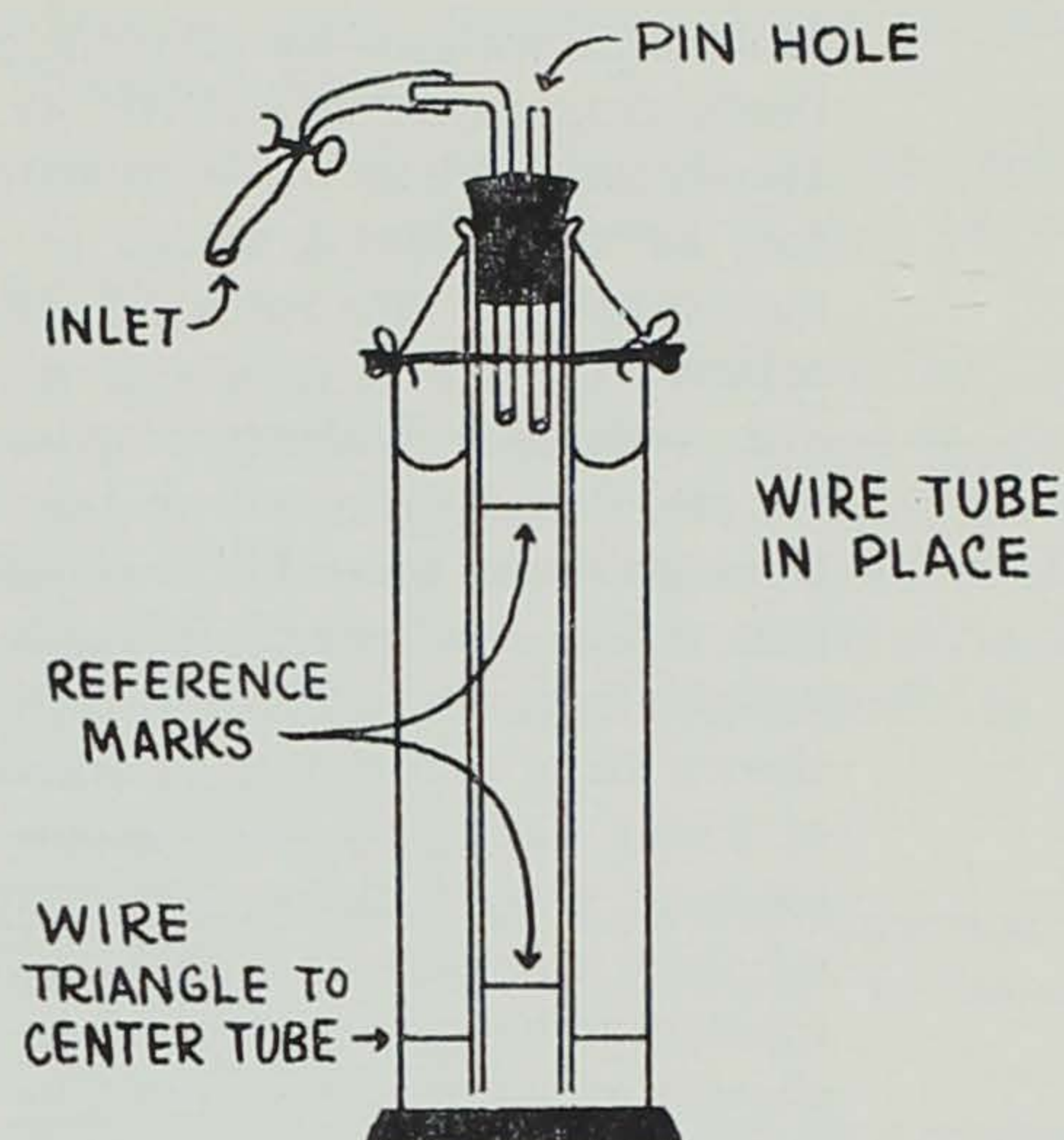


Figure 4

It is satisfactory for demonstration but not for the laboratory experiment. A better orifice is made by puncturing a piece of aluminum or platinum foil with a fine needle so as to leave no burr around the hole. The foil is then sealed to the end of the 6 mm glass tube, the aluminum with sealing wax, the platinum by direct fusion to the glass. Reference marks can be scratched on the outside of the inner tube with a file.

Fill the outer cylinder with water to above the top reference mark on the inner tube. Rinse the air from the inner tube by partially filling it with the gas to be used and allowing the gas to escape. Finally fill the tube with gas to below the bottom mark. Close the inlet tube with a pinch cock and allow the gas to escape through the pin hole. Measure with a stopwatch the time for the water level to move from the lower mark to the upper one. Repeat with each gas to be tested.

NOTE: Since rate of effusion is inversely proportional to the time of effusion of a fixed volume of gas, the relative *times* of effusion of equal volumes of gases at the same temperature are directly proportional to the square roots of their densities or molecular weights.

Molecular weights by effusion. Schilling's Apparatus (FIGURE 4). The apparatus used in the demonstration can be used for a series of pupil measurements on rates of effusion relative to that of oxygen as a standard. Molecular weights of gases can be calculated quite accurately from Graham's law provided the gases do not differ too much in molecular weight from that of water. Hydrogen shows large errors due to admixture of water vapor; so does carbon dioxide even though the water is saturated with this gas. Nitrogen and low molecular weight hydrocarbon gases show quite accurate results. The average molecular weight of natural gas can be determined if this is available.

Film: *Gas Laws and Their Application*. S.U.I. No. U-3968, 16 min., b & w. Sound. Or I.S.U. NS-3730, 13 min., b & w.

Film: *Molecular Theory of Matter*. S.U.I. No. U-298, 11 min., b & w. Sound.

Filmstrip: *Gases*. S.U.I. No. S 303.

- (8) Henry's law states that the solubility of a gas in a liquid is directly proportional to the pressure of the gas above the liquid.
-

Show how a liquid, such as a soft drink, that has been saturated with carbon dioxide under pressure gives up its dissolved gas at room pressure. Cool a bottle of pop, uncap it and attach a tested toy balloon. Allow the pop to return to room temperature as class continues or hasten the process by warming. The balloon will be blown up as the gas is liberated.

Problem work on the gas laws is suggested. Review Boyle's, Charles', and Dalton's laws.

D. The liquid state.

Film: *Mechanics of Liquids*. S.U.I. No. U-2856, 11 min., b & w. Sound.

1. In liquids the forces between molecules tending to hold them together are great enough to overcome the average kinetic energy of the molecules which tends to in-

crease the distance between the molecules and cause them to fly apart.

"Van der Waal's forces" is a term used for the molecular attractive forces which cause abnormalities in gases and liquids.

2. Kinetic theory applied to liquids.

- a. Liquids are not easily compressed because of the small space between the molecules.
-

Rapidly remove the heat from a small volume of water on a watch glass by placing it on a slab of dry ice. Lead the pupils to an understanding of the process of heat transfer based on the kinetic theory of gases and liquids.

Place a crystal of potassium permanganate in the bottom of a tall cylinder such as a 500 ml graduate filled with water. Cover with a watch glass and allow to stand. Have the class observe the process of diffusion of the KMnO_4 as it takes place over a period of several days or weeks. Lead a discussion of this phenomenon.

Make a careful layering of salt and fresh water in an 800 ml beaker or similar vessel. Place an egg carefully in the water. Promote a discussion of the results of this experiment.

- b. Liquids will expand because adding heat increases kinetic energy and overcomes intermolecular forces.
 - c. Liquids will diffuse because space still remains between the molecules, and their thermal motion causes mixing.
 - d. Liquids can be changed to solids by reducing the kinetic energy by removing heat energy. When this is done, the attractive forces overcome the random movement of the molecules.
 - e. Liquids transmit pressure because the molecules are free to move with respect to each other.
-

While hydrostatics and hydrodynamics properly belong in physics, chemists use many of these principles, such as that of the siphon, in connection with their laboratory work. Have pupils perform enough simple experiments and explain the principles involved so that it is certain they understand them thoroughly.

3. Surface tension is a force acting on the surface of a liquid due to the unequal forces of cohesion above and beneath the surface.
- — — —

Show a meniscus at the surface of a liquid in a narrow tube and explain why it occurs. Compare the meniscus of water with that of mercury. Discuss the difference between adhesion and cohesion and how the relation of these effects result in the difference in the menisci observed.

"Float" a razor blade on the surface of water. Why does the iron not sink?

Make a small wooden boat in the shape shown in FIGURE 5. Place a piece of gum

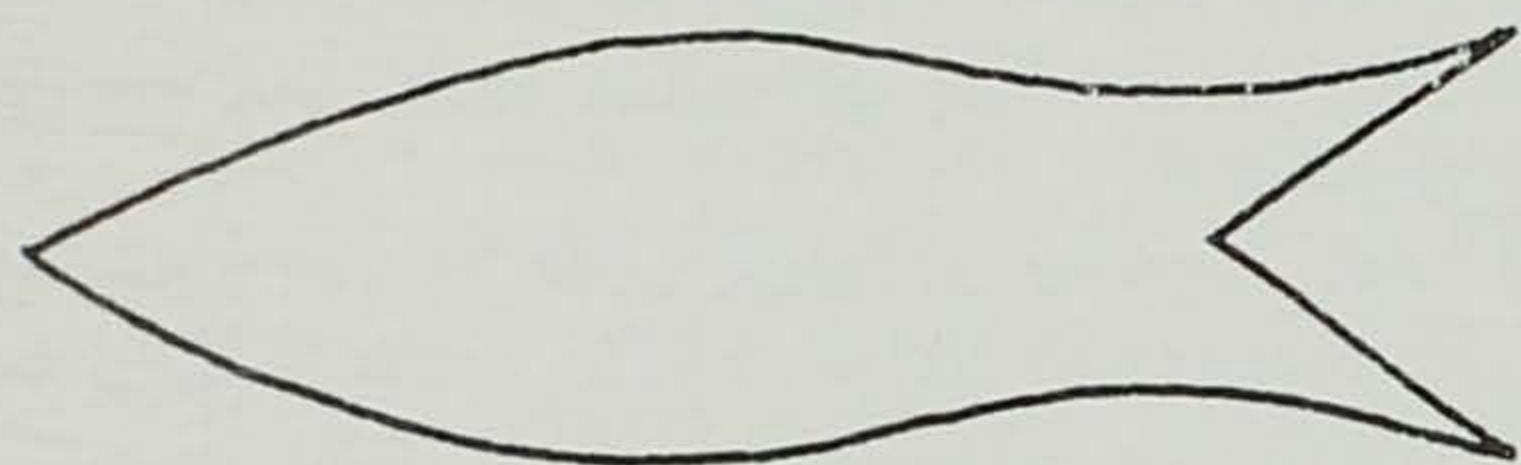


Figure 5

camphor in the V-shaped opening at the stern. Float the boat in a large vessel of water and observe the result. Have the class speculate on the reason for the behavior observed. Lead them to the correct conclusion, if necessary, by the observation that dissolved camphor lowers the surface tension of the water.

A glass filled with water above the rim being held in by surface tension will float a cork in its center. In a glass only partly filled with water the cork will be pulled to the glass.

Dip a common tea strainer or small metal sieve in melted paraffin. Shake off all excess wax and allow to cool. Later in class demonstrate that, due to the surface tension of water, you can carry water in a sieve.

Obtain a small tuft of natural cotton (NOT ABSORBENT COTTON). Weight it with a small wood screw and float the oily bit of cotton on the surface of water in a beaker. Add a trace of wetting agent such as Dreft or Alconox. The cotton will slowly lose its natural oil, be wet by the water, and sink. Tell the story of the duck that nearly drowned when wetting agent lowered the surface tension of the water in which it was swimming. See report of the Swampscott,

Massachusetts, meeting of the American Chemical Society.¹

Have pupils determine the surface tension of several liquids relative to that of water by the capillary rise method.

If a du Nouy ring apparatus is available, have them determine the surface tension of various concentrations of detergents plotting S. T. against the molar concentration of the solution. Have them draw conclusions as to the amount of detergent necessary to produce cleaning.

Suggested Project.

Have pupils devise a simple apparatus for determination of surface tension by the drop weight method.²

Have pupils investigate the determination of surface tension by the sessile bubble method.³

— — — —

4. Viscosity is a type of friction observed in fluids and is the result of molecular attraction. The attractive forces are not enough to keep the substance from flowing but are strong enough to give significant resistance to flow.
- — — —

Let two boys with egg beaters beat oils of different viscosities. Note the difference in their ability to sustain effort in heavy oil and in light oil. Explain this in terms of viscosity. If possible, obtain a set of falling ball viscosity tubes often available from a local oil company or filling station. Demonstrate the various viscosities of the oils they contain.

— — — —

5. Change of state is the change from gas to liquid, or from liquid to solid or vice versa.
- — — —

A simple cloud chamber can be made from a gallon jug fitted with a one-hole stopper and a short piece of glass tubing. Blow into the jug through the glass tubing to increase pressure. Put finger over end of tube and pull stopper. A small amount of water in the

¹ Caryl and Ericks, *Journal of Industrial and Engineering Chemistry*, XXXI (1939), 44; Howe, H. E. *Barron's Magazine*, XVIII (1938), 3.

² Morgan, *Journal of American Chemical Society*, XXX (1908), 360; and also Harkins and Brown, *Journal of American Chemical Society*, XL (1919), 519.

³ Tarter and Lingafelter, *Journal of American Chemical Society*, LXVII (1945), 2115.

bottom of the tube will cause nothing to happen in this case; but if you burn a match and drop it into the jug while burning and repeat the pressurizing, a cloud will be produced the next time.

Have pupils sublime iodine from an evaporating dish onto the underside of a Florence flask filled with cold water. Benzoic acid also can be used. Discuss with them what happens during a change of state.

— — — — —

- a. In condensation molecules move less rapidly as the average kinetic energy is decreased and Vander Waal's forces increase as distance between molecules is reduced.
- b. In evaporation the average kinetic energy of the molecules is increased. The distance between the molecules is increased until the kinetic energy overcomes Vander Waal's forces and the change to the gaseous state is accomplished.

— — — — —

Place a copper penny on a glass slide on microprojector. Put silver nitrate solution around penny and watch silver crystals form on screen.¹

Place alcohol on a pupil's skin. Have him explain what is happening in terms of kinetic theory and evaporation.

— — — — —

- c. Evaporation in a closed container results in the development of a pressure called vapor pressure.
6. Equilibrium is a situation where there is an apparent static condition but where, in reality, there is considerable activity.
7. A reverse reaction will occur if the products of a forward reaction are not allowed to leave the field.
 - a. In liquids, altering the concentration of the vapor above the surface can increase or decrease the rate in one direction or the other.
 - b. LeChâtelier's principle applies to vapor-liquid equilibria.
- E. The solid state.
 1. The attractive forces in solids are high.
 2. Kinetic theory applied to solids.
 - a. Attractive forces between molecules are greater than kinetic energy of molecules.
 - b. Distance is small between molecules.

¹ Y. Komuro and K. Sone, "Paper Made Crystal Models," *Journal of Chemical Education*, XXXVIII (1961) 580

- c. Diffusion occurs but is limited by small space.
- d. Change of state is accomplished by increasing kinetic energy of molecules.

— — — — —

Make a mold of a spoon in plaster of Paris and fill with molten Wood's metal. The spoon will melt in hot water, coffee, or tea. Save the mold to recast the spoon as a part of the demonstration.

Mix sand and lead oxide. Put mixture in test tube with water and mineral oil. Shake well. Note separation on basis of density.

— — — — —

- e. Density is related to mass of molecules and distance between molecules.
 3. The structures of solids take the shape of crystals.

A regular arrangement of particles with respect to shape, angle, surface, and edge.
 4. There are three kinds of crystals.
 - a. Metallic crystals.
 - b. Ionic crystals.
 - c. Macromolecular crystals.
 5. Faces represent possible arrangement of molecules and hence crystalline structure.
- Filmstrip: *Crystals*. S.U.I. No. 5830.
- a. Close packed.

Most simple twelve molecules touching interior sphere.
 - b. Face centered.
 - (1) Called cubic close packed.
 - (2) Has hexagonal arrangement.
 - c. Body centered.

A loose packing with only eight touching the inner sphere.
 6. Amorphous solids.

A peculiar condition related to bonding within a system wherein no crystals can be demonstrated.

UNIT II ATOMS AND MOLECULES (AN EXPANDED UNIT)

Electronic Structure of the Atoms

- A. Atomic structure²
 1. Dalton's Atomic Theory
 - a. Law of Definite Proportions
 - b. Law of Multiple Proportions

² Paul R. Frey, *College Chemistry* (Englewood Cliffs, New Jersey: Prentice-Hall, Inc., 1960), 83-107.

2. The electron.

a. Some general characteristics:

— — — — —
Lecture demonstrations on several topics in electrostatics, possibly employing an oscilloscope to demonstrate electrostatic and electromagnetic deflection of the electron beam.¹

NOTE: Molecular models are important in demonstrating structures and valence. Many kinds are available at a variety of prices. Atom models after Stuart and Briegleb distributed by La Pine Scientific Company, 6001 S. Knox Avenue, Chicago, Illinois, are excellent but expensive. Sanderson models are excellent, Extension Division, State University of Iowa. Many articles describing the construction of atomic and molecular models may be found in the *Journal of Chemical Education*. They are too numerous to list here. If construction is planned, several articles should be consulted and the one chosen which best matches the facilities available.

— — — — —

- (1) Move easily and faster by energy changes.
 - (2) They form an electronic field about the nucleus.
 - (3) Size of the atom is determined by its number.
 - (4) Electrons do not scatter atomic particles.
 - (5) Electrons move in spherical symmetry to the nucleus.
 - (6) When the energy of the electrons about the nucleus is at the lowest point, this is called "ground state."
 - (7) Mass of the electron.
 - (8) Charge of the electron.
 - (9) Speed of the electron.
- — — — —

Using Crooke's tube to show the electric discharge.

- (a) Maltese cross phenomena.
 - (b) Labeled diagram showing the tube's operational features.
- — — — —

3. The neutron.

a. Some general characteristics:

- (1) Size and mass of the neutron.
- (2) May decompose to a proton, electron, and neutrino.
- (3) Chadwick's studies in 1932.

4. The proton.

- a. Some general characteristics:
 - (1) Carries a positive charge.
 - (2) Conversion of matter to energy related.
 - (3) Positrons appear at the time of proton disintegration—matter to energy.

5. Nuclear structure.

- a. Atomic numbers—the number of protons in nucleus.
- b. Atomic weights.
 - (1) Meaning in terms of constituents.
 - (2) Isotopes.
 - (3) Evolution of standards.
 - (a) Hydrogen 1.000 system.
 - (b) Natural oxygen 16.000 system.
 - (c) Carbon isotope C¹² 12.0000 standard as of fall 1961.

6. Atomic spectra.

- a. Excitation of atoms results in changes of electron energy level.
 - (1) Evidence of energy changes from absorption and emission spectra.
- — — — —

Construct a simple spectroscope by using a slide projector, a screen and a glass prism. NOTE: Low cost replica gratings available as an alternate to a prism may be obtained from Edmund Scientific Company, Barrington, New Jersey.
Good for pupil projects on spectroscope construction.²

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- (a) Absorption lines in sun's spectrum-Fraunhofer.
- (b) Laboratory absorption spectra.
- (c) Emission (line) spectra of elements in visible region—Kirchhoff and Bunsen.
- (d) Niels Bohr and electronic configurations from visible and ultraviolet spectra.
 - i. Each electron with a definite energy.
 - ii. Absorption of energy—change in orbit or electron cloud

¹ Paul R. Frey, *College Chemistry* (Englewood Cliffs, New Jersey: Prentice-Hall, Inc., 1960), 103-104.

² Theodore A. Cutting, *Manual of Spectroscopy* (New York: Chemical Publishing Company), 1949.

c. Covalent bonds result in specific position of atoms within the molecules.

- (1) If moved far from this position the covalent bond no longer exists.
- (2) Bond strength is dependent on overlap of orbitals.
- (3) Covalent bonds are formed when neither atom has enough attraction to monopolize the electrons.
- (4) If the electronegatives are not far apart the possibilities are good for covalent bonding.
- (5) The smaller the atom, for a given oxidation state, the stronger the bond.

Exercises:

Supplementary Readings for Chemical Bond Approach, pp. 43-48—shows models for demonstrating electronegativity and partial charge.

Chemistry by Alfred B. Garrett, John S. Richardson, and Arthur S. Kiefer,¹ p. 106—shows a photograph with 2.2 million magnification of iron sulfide.

General Chemistry by Linus Pauling,² p. 203—shows electrolysis of mole in salt; p. 206—electrolysis of dilute aqueous salt solutions; p. 211—two electrolytic cells in series.

Laboratory:

Supplementary Readings for Chemical Bond Approach, pp. 125-139.

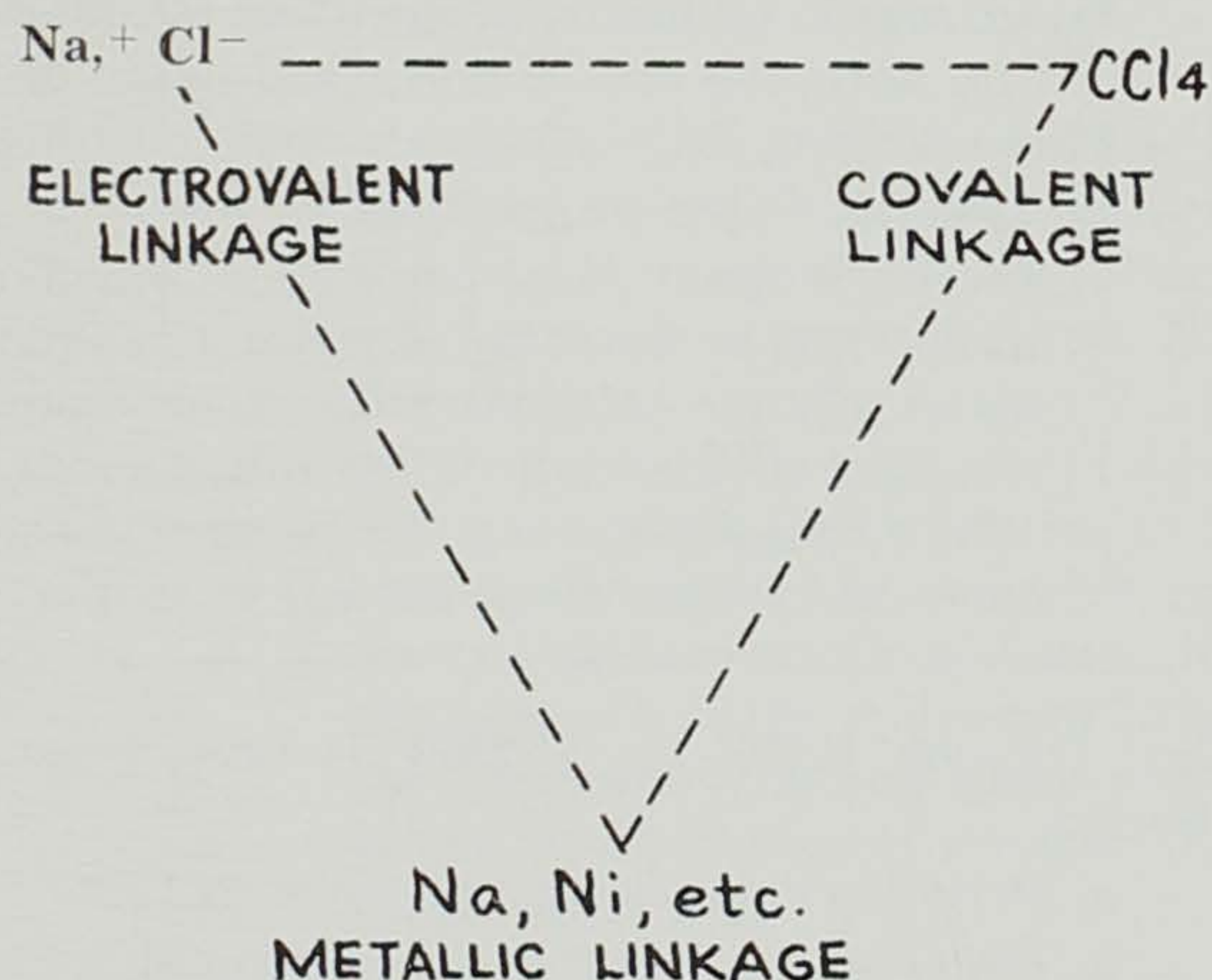


Figure 6

Make charts showing ionic bonding; see FIGURE 7.

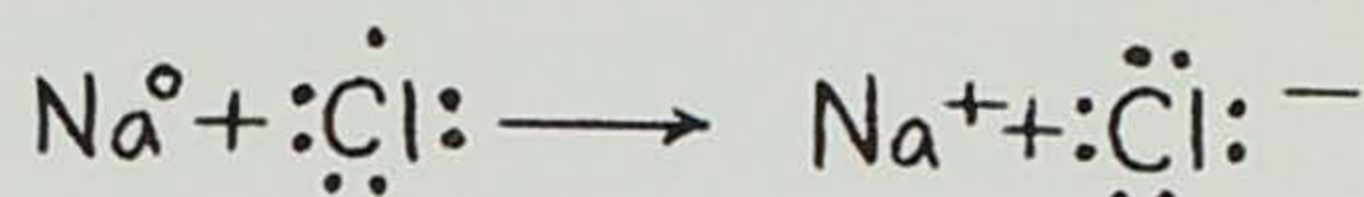
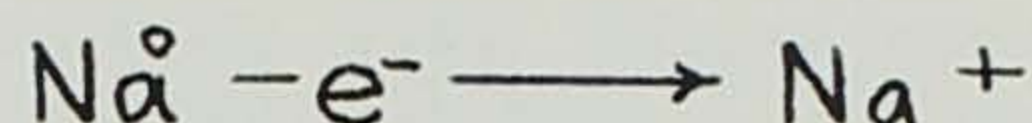


Figure 7

Prepare blank forms of "long" periodic table. Have pupils fill in symbols, atomic numbers, atomic weights, groups, and periods. (Do this up to and including Calcium.)

Models for Demonstrating Electronegativity and and Partial Charge—R. T. Sanderson in *Journal of Chemical Education*, XXIX (1952), 539; XXXI (1954), 2, 238; XXXII (1955), 140; XXXIII (1956), 443; XXXIV (1957), 195.

4. Periodic table and bonding.

- a. Group I A elements are strongly electropositive—release electrons easily.
 - (1) These form ionic compounds.
 - (2) Ions of this group are more stable than atoms.
 - (3) Ionic compounds are generally crystalline solids with a high melting point.
 - b. Group II A not as electropositive.
 - (1) Elements here tend to lose 2 electrons.
 - (2) Be and Mg are less reactive than others of this group.
 - (3) Large charge to atomic radius ratio results in strong bonds with water molecules and in high hydration numbers.
 - c. Group VII A most pronounced electronegativity.
 - (1) These are typical non-metals—these elements gain one electron.
 - (2) They are all strong oxidants—this emphasizes their toxicity, causing a marked oxidizing effect on living tissue.
 - d. Group VI A forms ions with a charge of minus two. Six valence electrons with a tendency to add two more.
A marked characteristic of the group is the allotropic modifications.
- #### 5. Energy effects accompanying chemical changes.
- a. General considerations. Chemical reactions are accompanied by energy effects

¹ A. B. Garrett, et al., *Chemistry* (Boston, Massachusetts: Ginn and Co.), 106.

² Linus Pauling, *General Chemistry* (San Francisco, California: W. H. Freeman and Co., 1953), 267.

which are usually large compared to those of physical changes such as mixing or simple solution.

- (1) The forming of chemical bonds results in the evolution of energy, usually in the form of heat but occasionally in the form of light or electricity, etc.
 - (2) The breaking of chemical bonds requires the absorption of energy usually in the form of heat, but occasionally in the form of light or electricity.
 - (3) The total energy change observed for a given reaction is the algebraic sum of all energy changes resulting from various steps in the process.
 - (4) A large evolution of energy* indicates a strong tendency for the reaction to take place as indicated by the chemical equation.
 - (5) A small evolution of energy* indicates that the reaction is not far from equilibrium.
- b. Thermal effects commonly observed.

Exothermic and endothermic effects on solution. Pour equal volumes of water into two beakers. To one add barium oxide, to the other add an equal amount of ammonium chloride. Barium oxide dissolves with the evolution of heat, ammonium chloride dissolves with the absorption of heat. Pass the beakers around the class as soon as possible so that pupils can observe the effects.¹

- (1) A reaction which gives off heat is called an exothermic reaction.
- (2) A reaction which absorbs heat is called an endothermic reaction.
- (3) Even though some reaction takes place with the absorption of heat, the overall change in energy* is that of energy being given off.
- (4) Heat effects accompanying chemical reactions are usually within the range of 5000 to 500,000 calories per mole of substance reacting as com-

¹ W. R. Carmody, "Dynamic Equilibrium: A Simple Quantitative Demonstration," *Journal of Chemical Education*, XXXVII (June, 1960), 312-313.

* Energy here is meant the "free energy" of the reaction as used in thermodynamics, which is the real measure of the tendency for a reaction to take place.

pared to 100 to 10,000 calories per mole for physical changes such as vaporization, crystallization, etc.

Films:

Atomic Alchemist (12 min., sd.)—1954 from Handel Film Corp., 6926 Melrose Ave., Hollywood 38, California.

Inside the Atom (10 min., sd.)—1948 National Film Board of Canada, 1270 Ave. of Americas, New York 20, New York.

The Electron (16 min., sd.)—Castle Film Div., United World Films, Inc., Russ Bldg., San Francisco 4, California.

Neutrons and the Heart of Matter (16 mm sound and color). Allied Chemical Corp., Nitrogen Division, 40 Rector St., New York 6, New York.

UNIT III

CHEMISTRY OF COMMON GASES

- A. Oxygen and its compounds—our most abundant element on earth.
1. Discovery and occurrence.

Draw attention to everyday uses of oxygen and show the effect of its pressure on the rates of combustion reactions. Collect several bottles of oxygen (not dried). Show the rapidity of oxidation with each of the following elements: sodium or potassium, powdered antimony, iron wire (use fine stranded picture wire; dip end in sulfur and ignite before plunging into oxygen), and burning phosphorus. Other metals may be tried.

The determination of per cent of oxygen in air may be made by absorbing oxygen in a solution of pyrogallic acid made strongly alkaline with sodium hydroxide. White phosphorus may also be used to absorb the oxygen from air but is slower.²

Film: *Oxygen*. S.U.I. No. U-2366, 11 min., b & w. Sound.

- a. Priestley's discovery and work in 1774.
- b. Occurrence—free in air, combined with metals and nonmetals of most kinds.
- c. Industrial sources.
 - (1) From liquid air.
 - (2) From electrolysis of water.

² Hubert N. Alyea, *Tested Demonstrations in General Chemistry* (Easton, Pennsylvania: Chemical Education Publishing Company, 1960), 23.

d. Laboratory preparation.

Oxygen may be prepared by the action of enzymes in yeast on 6 per cent solution of hydrogen peroxide.

Try potassium chromate as a catalyst in preparation of oxygen from potassium chlorate. Use this instead of the usual MnO_2 . Less is required.

Have pupils prepare oxygen in the laboratory by the addition of manganese dioxide to a 6 per cent hydrogen peroxide solution made by diluting the 30 per cent solution. They should collect the gas over water and test those of its properties as the instructor selects for emphasis. This method is safer than that of heating KClO_3 commonly used. If KClO_3 and MnO_2 are used, HAVE PUPILS MAKE CERTAIN THAT THE KClO_3 NEVER MAKES CONTACT WITH THE RUBBER STOPPER, NOT EVEN THE FINE FOAM OCCASIONALLY FORMED. THIS WILL CERTAINLY CAUSE AN EXPLOSION!

NOTE: 3 per cent H_2O_2 may be used, but it is slower than the more concentrated solution.

- (1) From metallic oxides.
- (2) From electrolysis of water.

Demonstrate electrolysis of water for the preparation of oxygen. Call attention to the ratio of the volumes of hydrogen and oxygen formed. Note that these will not be exactly 2:1 because of the difference in solubility of the two gases. Call attention to the sign of the electrode at which each gas is liberated and emphasize that electrons are added to liberate hydrogen (reduction) and electrons are removed in liberation of oxygen (oxidation of water).

- (3) From salts as KClO_3 .
- (4) From hydrogen peroxide, etc.

2. Properties and uses.

- a. Physical properties—allotropy.
- b. Chemical properties and uses.

Spontaneous ignition can be effected with

a half teaspoon of sodium peroxide placed on a two-inch cone of starch, sawdust or finely chopped paper. Lay a small chip of ice on the cone. Sufficient heat and oxygen will be released by the reaction of water from the ice with the sodium peroxide to ignite the material. Kindling temperature and oxidation can be discussed.

Dissolve a small amount of white phosphorus in carbon disulfide. A few drops of this solution on filter papers placed on top of a number of empty glass cylinders will illustrate: (1) spontaneous ignition as P catches fire, (2) oxidation and combustion as P catches fire and paper burns, (3) combustion and explosion as CS_2 and air explode in cylinder, and (4) incomplete combustion as evident by sulfur deposited on cylinder walls.

Fill a half gallon syrup can with natural gas having a hole cut on the side near the bottom and another in the center of the lid. Do not put the lid on tightly. Ignite the gas as it comes from the hole in the lid. Flame will at first be large and luminous. It will change to intensely hot flame as air is drawn in from bottom hole and mixed with gas. The gas and air mixture will eventually explode on reaching the proper proportions. Be careful and do not hurry the process.

Combustibility of certain dust particles in air can be vividly demonstrated by placing cornstarch in a handkerchief or cloth bag and dusting it through the cloth mesh into a flame.

A handkerchief saturated with 70 per cent alcohol diluted with an equal amount of water will not burn when the alcohol is ignited. Keep the handkerchief moving. Fill a beaker with carbon dioxide. Float a soap bubble on it. Call attention to properties of the oxide as opposed to pure oxygen.

Film: *Combustion*—Manufacturing Chemist's Association—Southerland Educational Films, Inc., 201 N. Occidental Blvd., Los Angeles 25, California.

Film: *Science of Fire*—General Electric Co., Schenectady, New York.

- (1) Rapid oxidation—combustion.

- (2) Slow oxidation—natural process—may result in spontaneous combustion.

- (3) Laboratory and industrial uses.

3. Compounds of oxygen.

Have pupils observe properties of the common metallic oxides, and of nonmetallic oxides. Have them test water solutions of each with litmus. Conclusions?

See UNIT X. A. 2. c. for experiments on catalysis which could be used here at the discretion of the instructor.

- a. Metallic oxides.
b. Nonmetallic oxides.
c. More complex compounds of some importance.

4. Bonding of oxygen.¹

- a. Principles of bonding—review these principles found under UNIT II.
b. Ionic bonding.

Distribute mimeographed copies of the electronegativity scale of the elements.²

Relate this to bond types obtained in compounds of elements of the periodic group. Stress the fact that while it is useful in predicting bond types, there are limitations to the dependence which can be placed upon these predictions.

- (1) Salt-like ionic compounds are formed between oxygen and metals of the alkali and alkaline earth groups.
(2) Ionic bonding is brought about by large differences in electronegativity between the two bonded atoms.

c. Covalent bonding.

If an ozonizer and cylinder of oxygen are available, demonstrate the odor and great oxidizing energy of ozone. If ozonizer is not available, see UNIT II. B. 2. for project on making an ozonizer. Explain during demonstration that the energy needed to form ozone from oxygen can be supplied from ultraviolet light as well as from electricity. Relate this to the use of

odor-destroying ultraviolet lamps as well as to the surface cracking of rubber tires. It is a good demonstration to attach a short length of fresh rubber tubing to an ozonizer while in operation. At the end of a few minutes the tube will show surface cracking and may even break off. CAUTION: OZONE IS POISONOUS—USE GOOD VENTILATION.

- (1) Oxygen molecule and ozone. Discuss the nonpolar nature of the oxygen molecule and its paramagnetic nature indicating unpaired electrons.
(2) Show bonding of ozone and mention nonmagnetic nature indicates paired electrons.
(3) Oxygen also exhibits coordinate covalent bonding in its compounds, e.g., hydronium ion.
(4) Bonding in many compounds is strengthened by resonance, e.g., in ozone.

B. Hydrogen and its compounds.

1. Discovery and occurrence.

Make mimeographed copies of the electronegativity scale and distribute them to the pupils. Use the periodic table to familiarize the pupil with the use of the electronegativity values. Stress the fact that while there are limitations to such use, they can be used to predict bond types.

- a. Hydrogen was prepared by Cavendish in 1766.

- b. Hydrogen exists mostly in combined form as water, coal, and petroleum products.

c. Industrial sources.

- (1) Prepared by passing steam over hot coke or hot iron.
(2) The reaction of steam with natural gas in the presence of a catalyst is an important source.
(3) Electrolysis of water provides a pure but expensive source of hydrogen.

d. Laboratory preparation.

- (1) Hydrogen may be prepared by the action of certain metals, usually zinc, on acids, usually hydrochloric acid.
(2) Hydrogen may be prepared by the action of more active metals on water.

¹ Marshall Sittig, "Oxygen, Chemical Raw Material," *Chemical and Engineering News*, XXXIX (1961), 93.

² Linus Pauling, *Nature of the Chemical Bond* (third edition; Ithaca, New York: Cornell University Press, 1960), 89-90.

Preparation and properties.¹

CAUTION: Hydrogen and oxygen for demonstration purposes may be conveniently supplied from high pressure cylinders, but these MUST BE ATTACHED TO SOME SOLID SUPPORT to prevent falling with the real danger of breaking off the outlet valve. The recoil from gas escaping at high pressure can cause great damage.

Hydrogen gas when mixed with air or oxygen in correct proportions is highly EXPLOSIVE. Do not allow mixtures of hydrogen with oxygen in large closed containers.

Have pupils attempt to prepare hydrogen by the addition of various metals to hydrochloric acid. Allow use of metals below as well as above hydrogen on the activity (electromotive) series. Have them collect hydrogen in a test tube by downward displacement of air and ignite this so as to obtain a "bark."

2. Physical properties of hydrogen.

a. Hydrogen is the lightest of all gases.

Demonstrate the rapid diffusion of hydrogen as compared to air.

Test a rubber balloon for leaks by filling with air and holding under water. If no bubbles are observed, allow the air to escape and fill the balloon with hydrogen.

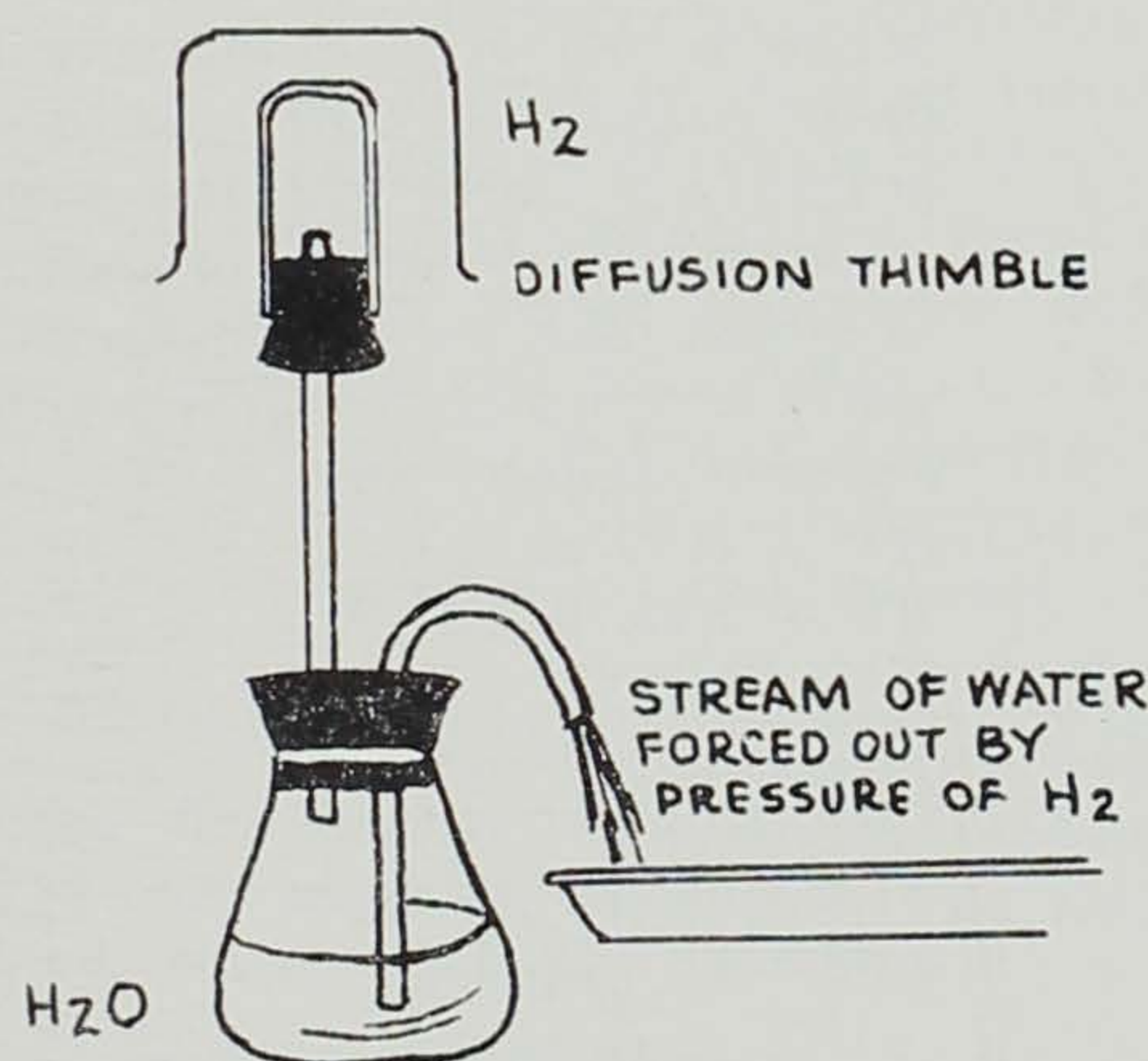


Figure 8

¹ Hubert N. Alyea, "Tested Demonstrations in General Chemistry," *Journal of Chemical Education*, XXXVII (1960), 9. (reprint)

The balloon will lose most of its hydrogen within a few hours due to diffusion of the gas through the rubber.

Demonstrate differential diffusion through a porous porcelain diffusion thimble in the apparatus shown in FIGURE 8.²

- b. It is colorless, odorless, and tasteless.
 - c. The low temperature (critical temperature—239.9° C) necessary to liquefy hydrogen indicates the small attraction of hydrogen molecules for each other.
- ## 3. Chemical properties of hydrogen.
- a. Hydrogen will combine directly with most elements under the proper conditions.

If a good hood is available, have pupils fill a 500 ml wide-mouth erlenmeyer flask with chlorine. See UNIT IX. A. 3. b. for preparation. Immediately lower a burning jet of hydrogen into the flask. The hydrogen will continue to burn in chlorine producing copious amounts of HCl which will fume in damp air and show the properties of an acid when dissolved in water.

- (1) Hydrogen will combine with the more negative elements with the evolution of large amounts of energy.

Explode a mixture of hydrogen and oxygen in soap bubbles. Bubble making solution may be prepared from household detergent and water, or bubble solution may usually be purchased at the local variety store.

NOTE: Alyea and some others explode the mixture in a "Coke" bottle. Use Caution! Tape the bottle with strong adhesive tape for safety.

- (2) Before hydrogen will react chemically, the H-H bond must be broken. H-H bond energy is approximately 103 kilo calories per mole, a comparatively high figure.
- b. Hydrogen also reacts directly with many compounds.
 - (1) The addition of hydrogen to certain molecules (hydrogenation of unsaturated vegetable oils, see UNIT XII.

² Hubert N. Alyea, "Hydrogen," *Journal of Chemical Education*, XXXVII (1960), A95.

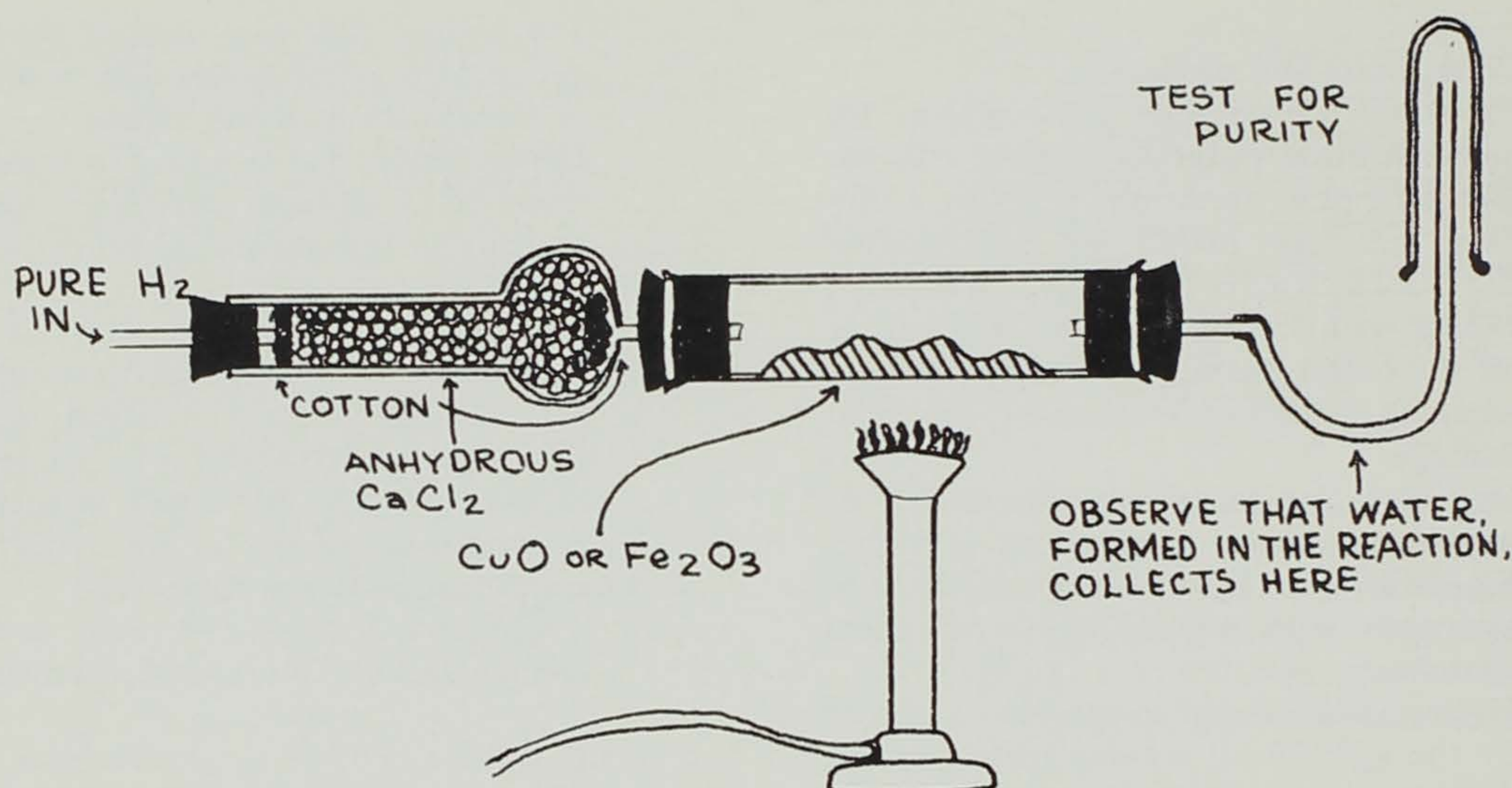


Figure 9

B. 1. g.) is of very great industrial importance. Crisco is an example of a solid fat produced from oils in this way.

- (2) Hydrogen also removes atoms from compounds. An example of this is the removal of oxygen (reduction) of copper oxide.

Better pupils may reduce CuO or Fe_2O_3 by passing pure dry hydrogen over the oxides heated in a Pyrex glass tube as shown in FIGURE 9.

Care! Test the purity of the hydrogen before applying heat to the tube by filling an inverted test tube held over the jet at the far end of the apparatus, removing it to a distance, and lighting it. A soft puff indicates pure gas escaping from the apparatus. A "bark" indicates oxygen present due to a leak or due to the fact that not enough oxygen has passed through to drive out the air.

4. Bonding in hydrogen.
 - a. Principles of bonding.
 - (1) Review the principles of bonding developed in UNIT II.
 - b. Ionic bonding of hydrogen.

Call attention to the fact that hydrogen

forms compounds which ionize in water or other polar solvents to form protons (hydrogen ions) which immediately become hydrated (solvated) but that H_2 dissolved in water does not show this behavior *except* when catalyzed by an active platinum surface. See UNIT XI. E. 2. on the hydrogen electrode.

- (1) Removal of an electron from a neutral hydrogen atom results in a positive charged ion or bare proton.
- (2) The proton is much too small and reactive to remain as an individual ion in the presence of other species. The proton will move into the electron cloud of nearby anions or move out to attach itself to such molecules as water or ammonia to form positively charged complex ions.

$$\text{H}^+ + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+$$

$$\text{H}^+ + \text{NH}_3 \rightleftharpoons \text{NH}_4^+$$

c. Covalent bonding of hydrogen.¹

- (1) The hydrogen atom will combine with another hydrogen atom in such a manner that electrons are shared.

Show the electron structure of the

¹ Linus Pauling, *The Nature of the Chemical Bond* (Ithaca, New York: Cornell University Press), 67.

covalent hydrogen molecule and of several hydrocarbons.

- — — — —
- (a) The atoms are identical and hence they share the electrons equally.
- (b) Equal sharing of the electron pair forms a nonpolar bond.
- (2) When hydrogen combines with other elements, the sharing is not equal. Unequal sharing of the electron pair forms a bond with partial ionic character.

— — — — —

Show the structure of HCl, HBr, HI, also of NH_3 , PH_3 , AsH_3 , SbH_3 showing the placement of electrons and the different bond angles. See Linus Pauling's *The Nature of the Chemical Bond*, p. 110 for bond angles, pp. 123 and 89-90 for scales of electronegativity. Call attention to the table of electronegativities distributed during the discussion of oxygen if UNIT III. A. 4. b. has been followed. Have pupils calculate from the ideal gas law the pressure in atmospheres necessary to compress one mole of H_2O as a vapor at room temperature to 0.018 liters, the volume of one mole of liquid water. Point out that the difference between this very high pressure and one atmosphere represents the internal attraction of water molecules, largely due to dipole-dipole attraction and to hydrogen bonding. If a good set of atomic models is available, it is worthwhile to spend a laboratory period assembling molecules of various compounds in which hydrogen is an important element. Suggested are the common acids, alkalis, hydrocarbons, and hydrides.

— — — — —

Examples of bonding of hydrogen of this type are the molecules of water, ammonia, and hydrogen chloride.

- d. Hydrogen bonding.
- (1) The sharing of a proton between two atoms is called hydrogen bonding.
- (2) Hydrogen bonds are generally formed only between small electronegative

elements as fluorine, oxygen, and nitrogen.

- (3) Hydrogen bonding is responsible for many of the unusual properties of water.

UNIT IV

FORMULAS AND EQUATIONS

A. Meaning and use of formulas.

1. Definition: A formula is a combination of the symbols of the constituent elements of a chemical compound.
2. Meaning and dual significance.
 - a. The numerical subscript of each symbol represents the number of atoms of that element entering into a unit quantity (molecule) of that compound.

— — — — —

H_2SO_4 . This indicates that two atoms of hydrogen will combine with one atom of sulfur to form the compound hydrogen sulfate (sulfuric acid).

- — — — —
- b. The numerical subscript of each symbol represents the number of atomic weights (moles) of that element entering into a gram formula weight of that compound.

— — — — —

Also indicated is that two gram atomic weights of hydrogen (2.016 g) will combine with one gram atomic weight of sulfur (32.064 g) and four atomic weights of oxygen (64.000 g) to give one gram molecular weight of hydrogen sulfate (98.080 g). The dual concept of formulas is important.

If a balance is available, the simplest formula of MgO or of CuS or similar compound may be determined. Weigh a half gram of Mg or a gram of Cu wire in a clean porcelain crucible. Heat the Mg in air (use cover)¹ or the Cu with an excess of sulfur (hood) until reaction occurs. Drive off the excess sulfur by heating. Desiccate the contents, cool, and reweigh. Have pupils follow good quantitative techniques. The gain in weight of the vessel and contents is that of the nonmetal entering into combination. From the original weight of the metal and the weight

¹ King and Caldwell, *Laboratory Experiments in General Chemistry* (second edition; New York: American Book Co., 1963), 29-30.

of the nonmetal, the empirical formula can be calculated.

3. Types of formulas.

- a. Simplest or empirical formulas state only the relative number of atoms in the compound.
-

Empirical formulas (simplest), e.g., NH_4SO_4 ammonium persulfate,* CH_3 ethane. These show the ratios of the atoms but not the number actually combining.

- b. Molecular formulas state the actual number of atoms (or gram atoms) contained in a molecule (or gram molecule, mole) of that compound.
-

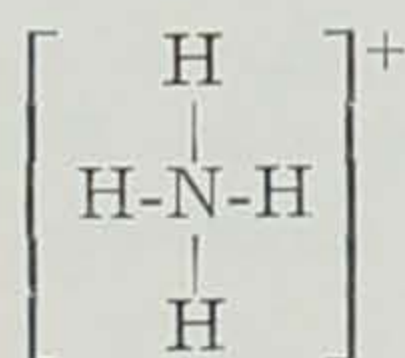
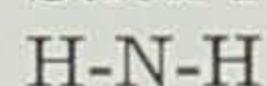
Molecular formulas (true formulas), e.g., $(\text{NH}_4)_2\text{S}_2\text{O}_8$ ammonium persulfate,* C_2H_6 ethane.

State the laws of Dulong and Petit and of Kopp. Show the method of obtaining molecular formulas of solid compounds from molecular heat capacities. Kopp's law states that molecular heat capacity (specific heat times molecular weight) is equal to the sum of the atomic heat capacities of the constituent elements. (Use also for A. 5. a.)

- c. Structural formulas attempt to indicate the relationship of the constituent atoms to each other in terms of their attachment or special relation.
-

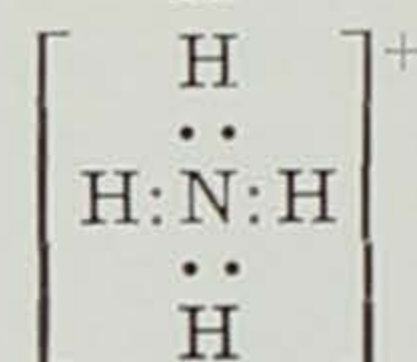
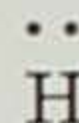
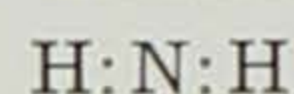
Structural formulas (indicate attachment of atoms).

Stick formulas



* Ammonium peroxy disulfate.

Electronic formulas



4. Uses.

- Formulas act as concise representations of the name of the compound.
- Formulas give the kind and number of atoms combined in the compound.
- Formulas allow the calculation of various quantities.
- Formulas allow the concise representation of chemical reactions by means of equations.
- Formulas of ions indicate the total external electrical charges (+ or -) of the ion from which the oxidation state of the various atoms may be inferred.
- Formulas of complex ions or compounds indicate the relationship of the various atoms or radicals.

5. Calculations based on formulas.

Molecular weights of substances in solution can be determined by freezing point lowering with fair accuracy if a thermometer reading in $1/10^\circ$ is available. Follow directions of the project, page 84. Urea in water or naphthalene in benzene work especially well. The benzene should be dried before using. Camphor is an especially good solvent where a precise thermometer is not available because of its high molecular freezing point lowering constant. Freezing point lowering constants for one mole of solute in 1,000 g solvent:

Solvent	K_f
Water	1.86 deg.
Acetic acid	3.90
Benzene	5.12
Bromoform	14.3
Cyclohexane	20.2
Camphor	40.0

$$\text{Mol. wt.} = \frac{K_f 1,000 w (\text{solute})}{\Delta T^\circ w (\text{solvent})}$$

NOTE: The determination of molecular weight by freezing point lowering can be postponed until colligative properties in

UNIT V.E.5. are studied. See also MOLECULAR WEIGHT BY FREEZING POINT DEPRESSION—PUPIL PROJECT.

- — — — —
- a. Calculation of formula weight of the compound or ion.

— — — — —

See A.3.b.

- — — — —
- b. Calculation of apparent molecular weight.
c. Calculation of percentage composition.

B. Nomenclature—inorganic.

1. Common or trivial names.
a. Used only for common compounds, e.g., water, sugar, etc.
b. Some are not scientifically correct, e.g., hypo for sodium thiosulfate.
2. Scientific or systematic names.

— — — — —

Board drill by pupils on the systematic naming of inorganic compounds is highly recommended.

Use of chemical handbooks is also recommended although the compounds referred to should be checked first by the instructor as the naming used in some handbooks does not always follow best modern style. Pupils may be given a list of formulas to name correctly. They may also be given the names and required to write the formulas.

Flat charts and rotary charts are available which are of some aid in teaching formulas and names, but their value as aids to learning is questionable. The instructor can easily satisfy himself as to their usefulness in his own situation by a short trial followed by a quiz.

Play Chemical Rummy.

CHEMICAL RUMMY

Practice in naming compounds can be obtained by marking 148 cards, four cards for each of the following common ions: Ag^+ , Al^{+++} , Ba^{++} , BO_3^- , Br^- , Acetate $^-$, Ca^{++} , Cl^- , Chlorate $^-$, $\text{CO}_3^{=}$, $\text{CrO}_4^{=}$, Cu^{++} , F^- , Fe^{++} , H^+ , HCO_3^- , Hg^+ , I^- , K^+ , Mg^{++} , MnO_4^- , Na^+ , NH_4^+ , NO_2^- , OH^- , $\text{PO}_4^{=}$, $\text{S}^{=}$, (O=), Sn^{++} , $\text{SO}_3^{=}$, $\text{SO}_4^{=}$, Zn^{++} , (others may be added). With these cards, several games may be played. One game is played by dealing each player five cards face down. One card from the deck is turned face up on the table to start the discard pile, the remaining cards are placed face down on the

table. The first player may pick up either the discard or draw to try to make a compound with what he has in his hand. When a compound is laid down, it must be named correctly by the player. If misnamed, corrector may claim compound. If corrector errs, a penalty of four points is assessed. On completion of each play the player must discard one card. Following players may pick up one to all the discards providing they use the bottom card picked up or they may choose to draw from the deck. Play until some one player holds no cards. Score ten points for going out first, five points for each compound put down, and subtract penalties plus the total valence of all cards held in hand. Game ends when someone has scored fifty or more points.

— — — — —

a. General principles.

- (1) Systematic names indicate the formula of the substance. Example: manganese dioxide is MnO_2 .
(2) Systematic names usually indicate the oxidation state of the key element. Example: Ferric trichloride or ferric chloride, FeCl_3 , indicates that iron is in its highest oxidation state of three.

b. Specific provision for correct naming.

- (1) Hydrogen or the metal involved is named first.
(2) Oxygen or other nonmetal is named last.
(3) Use of suffixes.
(a) -ide ending is used for binary compounds.
(b) -ate ending indicates ternary compound containing oxygen as well as two other elements.
(c) -ous ending indicates a lower oxidation state.
(d) -ic ending indicates a higher oxidation state.
(e) -ite ending used for salts derived from ternary acids, etc. Example: sodium hypochlorite, NaClO .
(4) Use of prefixes.
(a) Hypo- indicates a lower oxidation state than -ous alone. Hypochlorous, HClO .
(b) Per- indicates a higher oxidation state than -ic alone. Example:

perchloric acid, HClO_4 . Chlorine oxidation state is 7.

- (c) Greek numerical prefixes (mono, di, tri, etc.) give the number of atoms of each element present. Mono- is usually omitted.

(5) Other details to be introduced at instructor's discretion.

3. Organic chemistry has special Geneva nomenclature. See UNIT XII for basic rules. Also *Chemical Abstracts* for additional rules on organic nomenclature.

C. Equations.

1. General.

Board drill followed by homework assignments on equation balancing is highly recommended. Systematic methods of balancing should be stressed and required of the pupil.

- a. Equations indicate the substances used up and formed in chemical reactions.
b. Equations indicate the number of gram molecular weights (moles) used and produced in chemical reactions, therefore, can serve as a basis for calculations.
c. Equations indicate the relative volumes of gases used or formed in gaseous reactions.
d. Equations *rarely* give an indication of the actual mechanism of the reaction.

2. Types of equations.

- a. Combination (synthesis) $A + B = AB$.
b. Decomposition (analysis) $AB = A + B$.
c. Simple replacement $A + BC = AC + B$.
d. Double replacement $AB + CD = AD + CB$.

3. Specific information regarding equations.

Many equations should be balanced on the blackboard by the instructor. Both molecular and ionic types should be illustrated. The reasons for balancing equations should be reviewed.

- a. Equations must be "balanced" to be valid for use as a basis for calculations. This is because of the law of conservation of mass-energy.

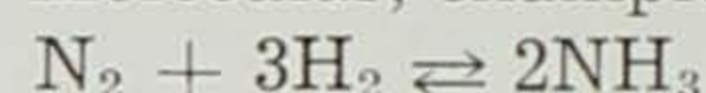
(1) In a balanced equation, the same number of atoms must be represented to left and to right of the equality sign or double arrows.

(2) In a balanced equation, the total mass of reactants represented must be equal to the total mass of products.

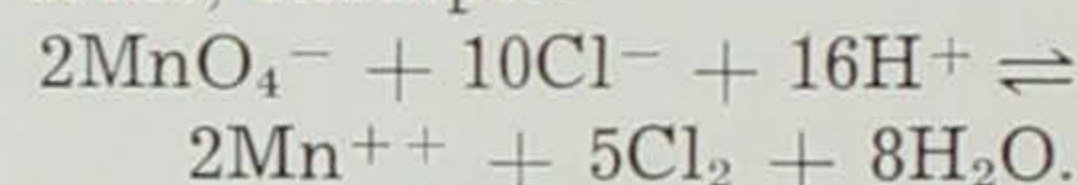
(3) In a balanced equation involving ions, the same net electrical charge must be represented on both sides of the equation.

b. Equations may be:

(1) Molecular, example:



(2) Ionic, example:



c. In a balanced equation, the number appearing before a formula (the coefficient) tells how many formula weights of that substance take part in the reaction.

d. Occasionally letters, (s), (l), (g), in parentheses following formulas give the physical form (solid, liquid, or gaseous) of the substances.

e. Some chemists use arrows pointing upward to indicate gases formed or downward to indicate precipitates.

f. Heat terms should not properly be included in chemical equations; rather a separate energy or heat equation should follow the chemical equation.

g. Balancing simple equations is usually done by observation, methods of balancing redox equations are given in UNIT VII. C.

D. Calculations based upon equations.

Problems of each type based on equations should be illustrated on the blackboard. Correct labeling of quantities should be stressed. The factor-label method of working more complex problems is a good one to teach.

Here is a good place to see that pupils really grasp the concept of the mole as a chemical entity. A good way to teach this is by using equations involving gases and showing the gas volumes at S.T.P. represented by the various formulas in the equation.

Emphasize the value of thinking in terms of *chemical* quantities, i.e., *moles*, in working chemical problems, rather than in terms of *grams*.

After teaching the method of determining the formula of a compound from its percentage composition, give the pupils the compound containing boron 3.6 per cent, uranium 78.9 per cent, nitrogen 4.6 per cent, and potassium 12.9

per cent and have them calculate the empirical formula. See who gets the answer first.

-
1. Calculations of moles of reactant to give moles of product or vice versa.
 2. Calculations of weight of reactant to give weight of product or vice versa.
 3. Calculations of weight of reactant to give moles of reactant or vice versa.
 4. Calculations of weight or moles of reactant to give volumes (S.T.P.) of gaseous products or vice versa.

Use the decomposition of potassium chlorate to determine accuracy of theoretic yield of a chemical equation.

-
5. Variations of above at discretion of the instructor.

Use gram-atom of magnesium to check weight-volume relation. Use HCl for non-oxidizing acid in this reaction.

UNIT V

WATER AND SOLUTIONS

Water is extraordinary in its behavior and structure. Many chemical reactions are profoundly influenced in speed and character by the presence of water.

A. The physical properties of water.

Call attention to the fact that although water has the same approximate density as other liquids, its properties are quite different.

Call attention to the fact that water is used as a standard to measure many things. Specific gravity and density should be discussed here.

-
1. Peculiar properties of water.

Molecules of water have spaces between them; this becomes evident when a long test tube or graduate is filled three-fourths full of water and is then completely and carefully filled to capacity with alcohol. Place palm over the end of container and invert. The container will no longer be filled to capacity after the liquids are mixed even though none is lost.

The wetting property of water is increased

by the addition of detergents. . . . Fill two containers with water. Add a detergent to one. Place a piece of wool yarn or non-absorbent cotton on each surface and observe time it takes for them to sink.

Have pupils determine the density of water.¹

-
- a. It will float its own solid form.

Some properties of water make interesting conversation. Demonstrate that ice is lighter than water by placing a large icicle in a milk bottle. Ice cubes may be used. Add cold water to fill the bottle while holding the ice under water. Let ice float and observe what happens as the ice melts.

-
- (1) Rivers and lakes do not freeze from the bottom nor usually become frozen solid, hence they thaw easily.
 - (2) Marine life becomes possible in cold climates.

- b. High specific heat.

Boil water in a paper cup. The paper will not burn until the water has boiled away. Why?

-
- (1) Used in defining the calorie.
 - (2) Makes water an efficient cooling agent —puts out fires.
 - (3) Has important effects on climate.

- c. High heat of vaporization.

Have pupils set up a typical distillation apparatus consisting of distilling flask, condenser, and receiver. Distill water from a solution of a colored, inorganic salt such as potassium permanganate or copper sulfate.

The following experiment demonstrates in an interesting fashion the heat of condensation (reverse of heat of vaporization) of water as well as several other phenomena connected with solutions.

Set up the apparatus shown in FIGURE 10. Boil the water in the first vessel, passing steam into the second until both

¹ L. E. Young and W. M. Petty, *Chemistry for Progress Laboratory Manual* (Englewood Cliffs, New Jersey: Prentice-Hall Inc., 1959), No. 8, 17.

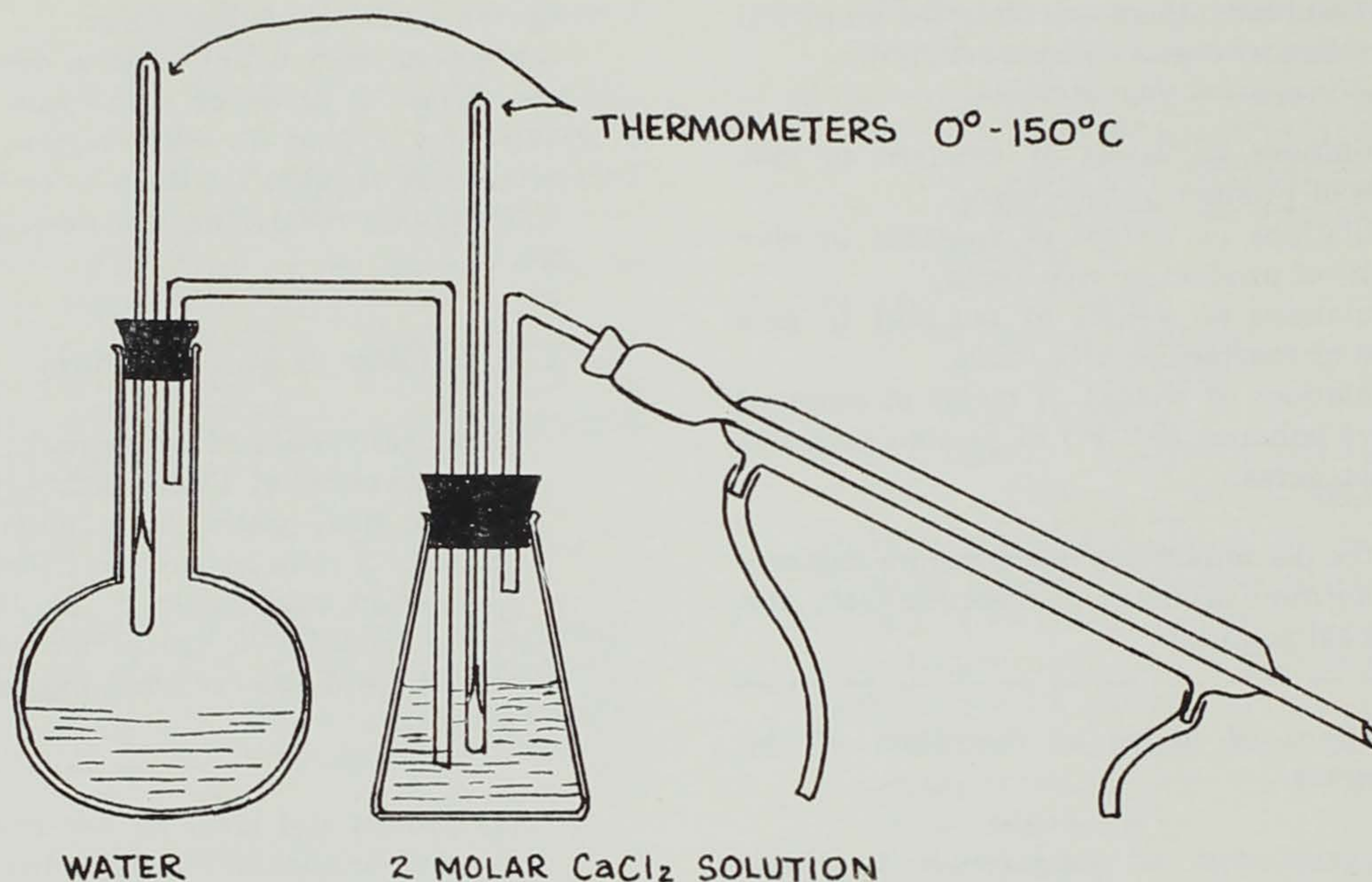


Figure 10

thermometers remain at constant reading. Read both thermometers and the barometer and explain the resultant observations.

-
- (1) Makes steam useful for heating buildings.
 - (2) Influences other uses for steam.
2. Comparison of properties of water with those of similar compounds.
- | Compounds | HF | H ₂ O | H ₂ S | H ₃ N(NH ₃) | CH ₄ |
|--------------|------|------------------|------------------|------------------------------------|-----------------|
| Mol. Wt. | 20 | 18 | 34 | 17 | 16 |
| Melting pt. | 83 | 0.0 | -82.9 | -78 | -184 |
| Boiling pt. | 9.5 | 100 | -61.8 | -33.4 | -161.5 |
| Heat of vap. | | | | | |
| cal./mole | 7200 | 9720 | 4485 | 5570 | 2208 |
3. Structural reasons for extraordinary properties.

Discuss the nature of a polar substance. Suggest other polar structures. Explain how molecular substances can be polar.

Make several sketches representing the water molecule. Stress that each sketch has its limitation but is valuable in that it aids in giving a mental picture.

Review the bonding energy which is responsible for the stability of chemical compounds. Suggest how the dielectric constant

of water can affect this. Discuss the nature of the electrostatic attraction involved in a polar molecule. Discuss the hydrogen bond and how it can produce a polymer structure in water.

-
- a. Electron structure of water.
 - (1) Tetrahedral arrangement of electron pairs as with NH₃ and HF.
 - (2) Two unshared pairs allow hydrogen bonding in three dimensions.
 - b. Spatial arrangement.
 - (1) Unsymmetrical arrangement of atoms gives polar properties.
 - (2) Results in closer grouping around hydrated cations (negative ends in) than around anions—cations thus more highly hydrated.
 - c. Polar properties.
 - (1) Results in associations of molecules in liquid water.
 - (2) Extraordinary solvent effects.
 - d. Hydrogen bonding.
 - (1) Responsible for high melting point and high boiling point.
 - (2) Results in other special properties.
4. Solvent properties.

Most substances, even glass, are soluble in

water. Grind a small amount of soft glass very fine and put it in a test tube with about 5 ml of distilled water. Add phenolphthalein, and note that the glass must be dissolving to produce the observed effect.

Most solids are more soluble in hot water than in cold water. Calcium acetate is a common exception to this. It shows negative solubility when heated and is precipitated when its water solution is heated.

-
- a. Dissolves small amounts of nearly all substances.
 - b. Especially good solvent for ionic solutes.
 - c. Poor solvent for nonpolar solutes.
- (1) Structural reasons for this.

B. Chemical properties of water.

Previous to the class exercise, prepare cobalt chloride paper by soaking sheets of good grade filter paper in nearly saturated cobalt chloride solution. Store these in a desiccator over anhydrous CaCl_2 . In class, pass out strips of this paper to each pupil and have each hold the paper tightly in the palm of his hand. After noting changes in color, attempt to arrive at an explanation through class discussion.

-
- 1. Neutral character (in general).
 - 2. Relatively stable with respect to chemical decomposition.
 - 3. Both weakly acidic and weakly basic.
 - a. Hydrolysis of salts.
 - b. Hydrolysis of esters—see UNIT X.B.2.b.
 - 4. A catalyst for many reactions.
-

Aluminum powder mixed with iodine will ignite when a drop of water is added.

-
- 5. Combines weakly with ions to form hydrates—see UNIT VI.B.8.
-

Pupils may experience heat of hydration by placing a small amount of CaCl_2 on their hand and adding a drop of water. The heat evolved is large and may be uncomfortable. Have water handy to cool hands.

-
- 6. Has oxidizing properties—see UNIT IX. E., F., and G.

C. Uses of water—our most plentiful bulk chemical in liquid form.

- 1. Importance and use as a chemical reagent.
-

To demonstrate how water aids chemical reactions, add a gram of potassium bitartrate and a gram of sodium bicarbonate to a test tube. Shake and note absence of reaction. Add a little water and observe effect. Silver nitrate and sodium chloride may also be used in a similar manner.

-
- 2. Physical-chemical uses.

- a. Industrial and sanitary uses.
- b. Potable water.
 - (1) Purification for drinking and food preparation.
 - (2) Softening of water.
 - (3) Desalting of sea water.

D. Conservation of water resources.

E. Solutions in general but with emphasis on aqueous solutions.¹

- 1. Terms relating to solutions.²
-

Demonstrate the selectivity of solvents using iodine as solute and water, alcohol, and carbon tetrachloride as solvents. Have pupils memorize the general rules of solubility, emphasizing effects of polarity of the solvent.

-
- a. Solvent.
 - b. Solute.
 - c. Saturation—a condition of equilibrium between free and dissolved solute.
 - d. Supersaturation—a nonequilibrium condition.
-

Demonstrate supersaturation by dissolving 50 grams of sodium acetate in 10 ml of hot water using a water bath to maintain the temperature. Remove the test tube and allow it to cool. Then add a small crystal of sodium acetate as a "seed."

Discuss the tendency of substances to crystallize in pure form from a mixture and the use of recrystallization as a method of purification of solids.

¹ Alan Holden and Phyllis Singer, *Crystals and Crystal Growing* (Columbus 16, Ohio: Wesleyan University Press), Chapter II.

² Manufacturing Chemist's Association, Inc., "Making a Solubility Curve—Experiment No. 8," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 15.

2. Factors aiding solution.

Give each pupil small amounts of solute, a beaker, water, stirring rod, burner, and strip of paper. Use a slow dissolving solute such as copper sulfate or a fast dissolving one which is easily traced such as potassium permanganate. Let each investigate various methods of aiding solution: (1) stirring, (2) convection (by placing solute in a loop of paper suspended just below the water surface), (3) rise in temperature, and (4) reducing the solute to fine powder before adding it to the solvent.

- Increasing the amount of solvent.
- Raising the temperature (except in a few cases).
- Decreasing the size of solute particles.
- Agitation of the system.
- Increasing pressure in the case of gaseous solutes.

3. Methods of expressing concentration defined.

Spend considerable class time on methods of expressing the concentration of solutions, especially the "chemical" methods of expressing this; i.e., molal, molar, and normal systems. Call attention to the effect of varying temperature on the concentration of solutions expressed in the molar and normal as opposed to the molal system.

- Per cent by weight.
 - Per cent by volume.
 - Molar or formal method.
 - Molal method—especially for non-aqueous solutions.
 - Normal method.
4. Colligative properties—effect of (normal) solutes on the physical properties of solutions which is proportional to the number of solute particles present per quantity of solvent.

Discuss how ionic substances effect freezing point, etc., differently than covalent substances.

Have pupils determine the molecular weight of a solute in solution by one of the colligative properties. Freezing point lowering is

best. For directions see page 84, Supplementary Lectures.¹

- Lowering of vapor pressure—Raoult's laws.
 - Rise in boiling point.
 - Lowering of freezing point.
 - Osmotic and diffusion phenomena.
5. Calculations involving solutions.
- Weight of solute to give a solution of a required concentration.
 - Volume of liquid solute as H_2SO_4 to give a solution of a required concentration.
 - Conversion of concentration expressed in one method to that in others, i.e., weight per cent to molar.
(NOTE: This as well as d. and e. may be curtailed or omitted at the instructor's discretion.)
 - Dilution problems.
 - Simple calculation of colligative properties as lowering of freezing point.

UNIT VI IONIZATION AND IONIC SOLUTIONS

A. Ions and their formation.

1. Since all atoms are electrical in nature, altering the ratios of negative to positive charges results in the formation of charged atoms called ions.

Make a Cottrell smoke precipitator. Attach one lead from a spark coil to a foil surrounding a glass tube of about one inch diameter. Extend a wire from the other terminal of the spark coil through the tube, insulated from the foil. Place a small amount of hydrochloric acid in one flask and some ammonium hydroxide in a second flask. With glass tubing, connect the flasks and large glass tubing in a train. Blow air into the first flask causing ammonium chloride to be forced into the Cottrell precipitator. Activate the spark coil and see smoke consumed. Explain the results of ionizing the smoke.

Ionization in a flame can be shown by holding a lighted match near a charged electro-

¹ Manufacturing Chemist's Association, Inc., "Finding Molecular Weights—Experiment No. 6," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 11; R. Mikulak and O. Rundquist, "Molecular Weights by Cryoscopy: A General Chemistry Laboratory Experiment," *Journal of Chemical Education*, XXXVIII (November, 1961), 557-558.

scope. Charged pith balls or balloons lose their charge rapidly when a flame is brought near.

2. Ions result from the transfer of electrons between atoms when reaction occurs.

Use pith balls and add electrons from a charged rubber rod. Relate the effect to what happens electrostatically within the chemical compound.

Use a static electricity generator such as a Van de Graaff or Wimhurst machine. Charge a sphere or condenser with varying amounts of electricity. Bring an electroscope close to it. Relate the effect to the activity of ionic substances.

- a. Atoms tend to reach the stable electron systems of the rare gases.
b. Atoms with nearly filled principal energy levels (shells) gain electrons and become negative ions or anions. Example:
 $\text{Cl}^0 + e \rightleftharpoons \text{Cl}^-$.

Aid the pupils in forming a mental picture of the nucleus with its shells of electrons and its role in ion formation. Be sure to discuss isoelectronic structure.

- c. Atoms with barely started shells lose electrons and become positive ions or cations. Example:
 $\text{Na}^0 \rightleftharpoons e + \text{Na}^+$.

3. Ions result from the dissociation of covalent compounds.

- a. Dissociation occurs in gases when sufficient electrical or other energy is available. Example:
 $\text{H}_2 + \text{energy} \rightleftharpoons \text{H}^+ + \text{H}^-$.

Discuss the difference between ionic and covalent compounds. Point out that there is a wedding of ionic and covalent bonding arrangements in many compounds.

- b. Dissociation occurs in solution when the covalent bond is weakened by highly polar solvent molecules. Examples:
 $\text{HCl (gas)} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{Cl}^-$
 $\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{CH}_3\text{COO}^-$.

Draw a picture of water as a dipole and explain its role in ionization.

B. Characteristics of ions.

Electrolyze a solution of potassium permanganate. Show the movement of the ions to the electrodes by placing a porous membrane to separate the electrodes. Watch the effects of the ionic migration.

- The chemical properties of an ion are different than those of the neutral atom from which it was formed.
- The charge on an ion depends upon the number of electrons gained or lost.
- The energies necessary to remove electrons from atoms to form cations are called ionization potentials.
 - The first ionization potential is the energy (measured in electron volts) necessary to remove the most loosely held electron from an isolated atom in the gaseous state.
 - The second ionization potential is the energy necessary to remove the second most loosely held electron under the above conditions.
- The energy released when an atom gains an electron to form an anion in the gaseous state is called the electron affinity.
 - Electron affinities are not directly measured as are ionization potentials. Absolute values are difficult to obtain.
 - Pauling has formulated a method of calculating relative electronegativity values which serve a similar purpose.
- Ions form bonds in salt crystals which are largely electrostatic (ionic) in character but which often have some covalent character.

The greater difference in Pauling's electronegativity values of the atoms involved, the more ionic in character will be the bond.
- The radius of an ion depends upon the number of shells of electrons in its structure.
- Cations tend to be smaller than anions because they are formed by loss rather than gain of electrons.
- All ions tend to be hydrated in water solution; cations more than anions.

C. Ionic solutes—acids and alkalis (bases).

1. Common properties of acids.
 - a. Sour taste.
 - b. Change colors of indicators (litmus to red).

— — — —

Indicators and pH. Arrange in pairs, ten glass cylinders and fill them about three-fourths full of water. Add to each pair one of the following: *thymolphthalein*, *phenolphthalein*, *phenol red*, *bromthymol blue*, and *methyl red*. Add to the thymolphthalein cylinders an amount of NH_3aq to produce the deep blue color of the indicator and add the same amount to each of the other cylinders. Break up dry ice to lumps of fairly uniform size. Drop into one cylinder of each pair and leaving the other for control. As carbon dioxide dissolves, the pH goes down. Demonstrator may have to add HCl to methyl red to get desired effect.

Have about 15 or 20 common household substances and have the pupils check them for acid-base content. Use litmus, hydron, methyl orange and phenolphthalein, or H-L universal indicator.

— — — — —

- c. Dissolve many metals with liberation of hydrogen.
 - d. Neutralize alkalis.
2. Formation of acids by:
 - a. Reaction of acid anhydride (nonmetallic oxides) with water, e.g.,
 $\text{SO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{H}_2\text{SO}_4$.
 - b. Treating a salt of the acid with H_2SO_4 .
3. Common properties of alkalis.
 - a. Soapy taste.
 - b. Change colors of indicators (litmus to blue).

— — — —

The patriotic colors are produced by pouring sodium hydroxide solution into three beakers containing one each of the following solutions: phenolphthalein, lead acetate, and copper sulfate. See C.1.b.

— — — — —

- c. Precipitate many insoluble hydroxides.
 - d. Neutralize acids.
4. Formation of alkalis by:
 - a. Reaction of basic anhydride (metallic oxides) with water, e.g.,
 $\text{CaO} + \text{H}_2\text{O} \rightleftharpoons \text{Ca}(\text{OH})_2$.

- b. Miscellaneous methods (at discretion of instructor).
5. Theories of acids and bases.
 - a. Arrhenius' theory—acids give H^+ , alkalis give OH^- to the solution.
 - b. Brönsted-Lowry theory—acids, proton donors; bases, proton acceptors.

— — — —

Be sure the pupils understand the concept of donor and acceptor.

Draw several equations on the board and explain how the proton is donated.

Call attention to the fact that Brönsted's theory is sometimes labeled under the heading of protolysis and explain how the term is an apt descriptive term.

— — — — —

- c. Lewis' theory—acids are electron pair acceptors, bases are electron pair donors.

— — — —

Call attention to the role of electron structure in Lewis' theory.

Bring out the idea of conjugate acids and bases. That is, an acid and a base combined to form an acid and a base of a different strength.

Call attention to the fact that Lewis' theory covers all possible circumstances surrounding acid or base characteristics.¹

— — — — —

- D. Ionic solutes—salts—substances with ionic lattices in the solid state.

1. Formation of salts.
 - a. By direct combination of the elements.
 - b. By evaporation to dryness of a solution resulting from the partial or complete neutralization of an acid by a base.
 - (1) Normal salts.
 - (2) Acid salts.
 - c. By evaporation to dryness or crystallization of a solution resulting from the solution of a metal or metallic oxide by an acid.
 - d. By precipitation from solution.

- E. Properties of ionic solutions.

1. Abnormal colligative properties. See UNIT V.C.4.

— — — —

Give the class some substances which will ionize. Have them measure the properties of

¹ John Baxter, "A Reaction to Introduce Lewis Acid-Base Chemistry," *Journal of Chemical Education*, XXXVIII (July, 1961), A491.

these substances in solution, their effect on boiling point, conductivity, and freezing point. See also UNIT V.E.4.

2. Conduction of electricity—distinction between electrolytes which do conduct and nonelectrolytes which do not.¹
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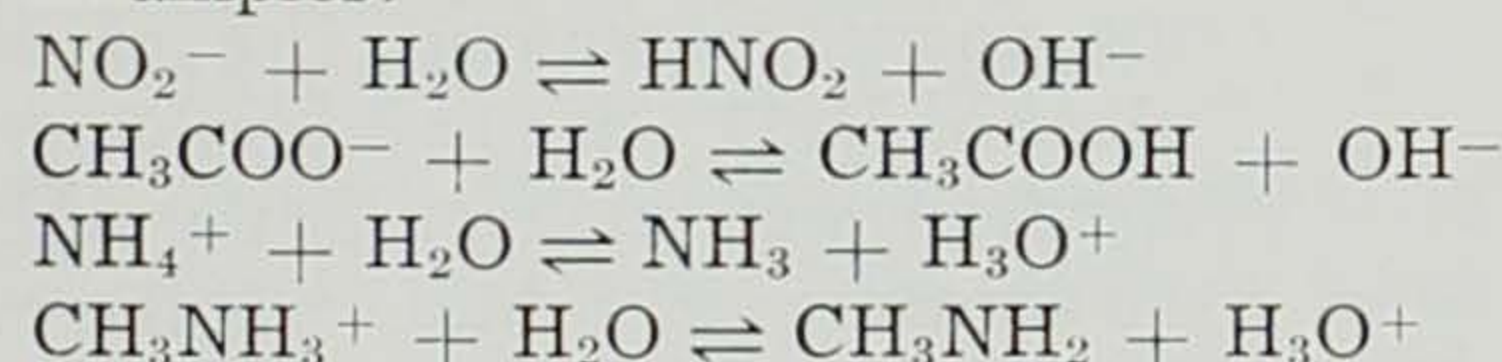
When explaining conductivity, ionization, and their relation, remember that hydrogen chloride does not show ionization in dry benzene nor does it conduct an electric current. In water, HCl becomes an excellent conductor.

For suggested equipment see project, "Electrolytic Conductivity Apparatus," page 79.

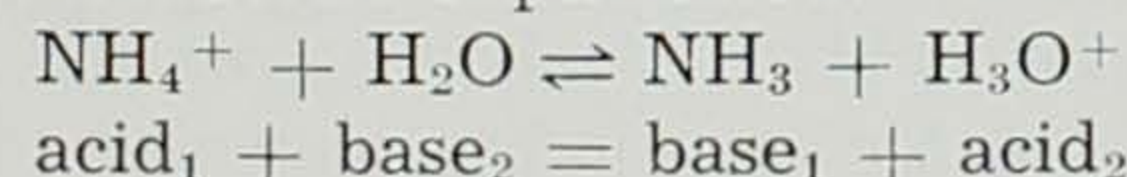
3. Degree or extent of ionization—distinction between strong and weak electrolytes.
- Strong electrolytes—salts and strong acids and alkalis which are completely or nearly completely in the form of ions.
 - Weak electrolytes—in which an equilibrium exists between molecules and ions (to be developed in detail in UNIT X).
 - Hydrolysis—the reaction of water with ions or molecules to form slightly ionized molecules.

NOTE: Brönsted looks upon hydrolysis as the competition of two Brönsted bases for the protons (H^+ ions) present in the solution.

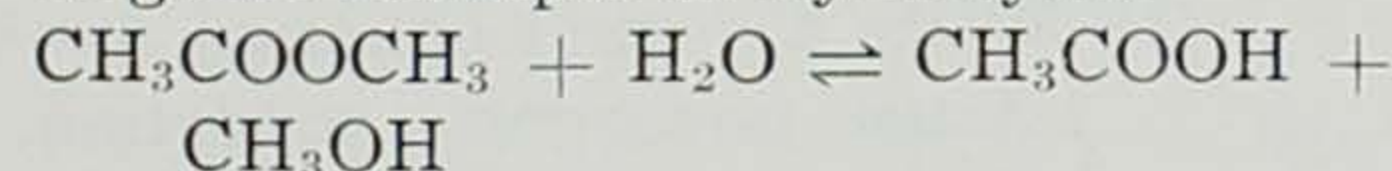
Examples:



Brönsted's explanation:



Organic example of hydrolysis:



F. Simple calculations involving ionic solutions.

- Methods of expressing hydrogen ion concentration.

¹ J. S. Proctor and J. E. Roberts, "Analysis of Aspirin: A Conductometric Titration," *Journal of Chemical Education*, XXXVIII (September, 1961), 471.

- Titrateable acidity—total available hydrogen ions (or other acidity) in moles per liter.
-

Work problems in pH using logarithms.

Find the pH of an acetic acid solution by first titrating to find its concentration and then calculating its pH using the known ionization constant in Ostwald's equation.

- Momentary acidity—concentration of hydrogen ions existing without additional ionization or reaction in moles per liter.
 - The concentration of hydrogen ions by simple calculations.
 - The pH scale of momentary acidity.
 - Calculation of pH from $[H_3O^+]$ and vice versa.
 - Estimation of pH by indicators.
-

Have pupils compare their calculated value with that obtained from indicators or from the pH meter made in project, page 71.

- Discuss the ranges of different indicators and suggest best for certain ranges.
- Control of pH buffers.
Example: $NH_3(aq) + NH_4^+$ as ammonia, ammonium chloride;
 $CH_3COOH + OAc^-$ as acetic acid, sodium acetate.

UNIT VII

OXIDATION-REDUCTION

A. Oxidation-reduction.

1. Oxidation.

Oxidation is any reaction that involves the loss of electrons.

2. Reduction.

Reduction is any reaction that involves a gain of electrons.

3. Redox reactions.

Place a large zinc strip in a solution of copper sulfate. Discuss what is being reduced and oxidized. Identify the reducing agent and oxidizing agents.

Rapid redox reaction.

Touch (CARE) a glass rod to which is attached a small crystal of $KMnO_4$ to a drop

of cold H_2SO_4 and touch the wick of a small alcohol lamp. The lamp should light.¹

- — — — —
- Oxidation and reduction must occur together.
 - An oxidation and reduction reaction is one in which there is a transfer of electrons.

4. Oxidizing and reducing agents.

Develop the concept of reducing agents as being electron donors and oxidizing agents as being electron acceptors.

B. Oxidation number.

1. Arbitrary number.

- Only in extremely electrovalent bonds does a complete transfer of electrons take place.
- The assignment of a definite oxidation number is useful in identifying the different oxidation states of the elements.

2. Rules.

- The oxidation number of a free element is 0.
- The oxidation number of hydrogen in all compounds except metal hydrides is +1; in metal hydrides it is -1.
- The oxidation number of oxygen in all simple compounds except peroxides is -2; in metal peroxides the oxidation number is -1.
- The oxidation number of a simple ion is equal to the charge on the ion.
- The algebraic sum of all the oxidation numbers of the elements that are combined in a radical is equal to the charge on the radical.
- The algebraic sum of all oxidation numbers in a compound is equal to zero.

C. Balancing oxidation and reduction equations.

1. Methods of balancing.

a. The valence change.

This method is useful when the reactions are nonionic in nature.

b. Ion-electron.

This method is most useful when the reactions are ionic in nature.²

c. The ion-electron method will be developed in this unit.

2. Balancing redox equation by the ion-electron method.

- A balanced equation must balance electrically and atomically. This means that the total charges on each side of the equation must be equal and the atoms on each of the equations must be equal.
- Essentially this method consists of dividing oxidation-reduction reactions into two parts or two half-equations. One-half equation represents an oxidation, while the other half-equation represents a reduction.
- The instructor will find it worthwhile to use only ionic oxidation-reduction equations at first. Pupils should be given ionic oxidation-reduction equations to balance. Once the pupil has mastered the technique of balancing by the ion-electron method, he can be introduced to the molecular form.

3. Rules for balancing.

Using the example given below, the ion-electron method will be developed. The instructor should use a simpler example at first before attempting to balance the more complicated example used in this unit. Example: $\text{H}^+ + \text{MnO}_4^- + \text{Fe}^{++} \rightarrow \text{Fe}^{+++} + \text{Mn}^{++} + \text{H}_2\text{O}$.

a. Step 1.

Use the oxidation numbers to identify the oxidizing and reducing agents. Only molecules or radicals entering into the reduction or oxidation should be used in the partial equations. Use of Mn^{+7} in place of MnO_4^- is not desirable since this species is not involved in the reaction. In the above reaction MnO_4^- is the oxidizing agent and the product is Mn^{++} . The reducing agent is Fe^{++} and the product is Fe^{+++} .

b. Step 2.

- Write a partial equation for the substance undergoing oxidation, the reducing agent. With strong acids and bases and with salts write down only the ions taking part in reaction. Be careful to show the electrical charge of the ion. Weak electrolytes and nonelectrolytes should be written as molecules.

¹ I. Fine, *Journal of Chemical Education*, 929 (1931), 8; Manufacturing Chemist's Association, Inc., "Oxidation-Reduction—Experiment No. 22," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 43.

² Oscar E. Lanford, *Journal of Chemical Education*, 929 (1931), 636-667.

Examples: CH_3COOH ; CH_3CHO ; H_2O .

- (2) Balance the partial equation in terms of atoms.

Example: $\text{Fe}^{++} \rightarrow \text{Fe}^{+++}$.

- (3) Balance the partial equation in terms of electrical charges. This is accomplished by showing the number of electrons lost. Show the loss of electrons on the right side of the equation. Example: $\text{Fe}^{++} \rightarrow \text{Fe}^{+++} + e^-$. The partial equation is balanced when the same net charge appears on both sides: i.e., $2+$ and $2+$. The partial equation is now balanced atomically and electrically.

NOTE: Water or OH^- ions may be used as products or reactants if the reaction takes place in an alkaline medium.

Example: $\text{CrO}_2^- + 4\text{OH}^- \rightarrow \text{CrO}_4^{=4} + 2\text{H}_2\text{O} + 3e^-$.

NOTE: In cases where oxygen is needed to form the product ion two OH^- ions will usually be needed for each O atom supplied.

c. Step 3.

- (1) Write the partial half-equation for the substance undergoing reduction, the oxidizing agent. With strong acids, bases, and with salts write down only the ions taking part in the reaction. Be careful to show the electrical charge of the ion. Weak electrolytes and nonelectrolytes should be written as molecules.

Example: $\text{MnO}_4^- \rightarrow \text{Mn}^{++}$.

- (2) Balance the partial half-equation in terms of atoms. This reaction takes place in an acid medium and so H^+ and H_2O may be supplied. When an ion or molecule which contains oxygen is reduced in an acid medium, the oxygen lost will usually go into water.

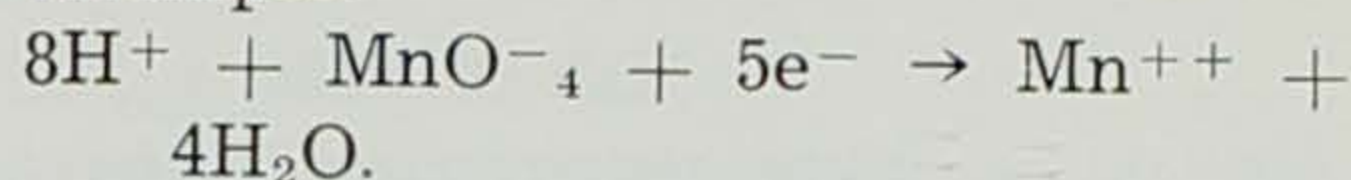
Example: $8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{++} + \text{H}_2\text{O}$.

d. Step 4.

Balance the partial equation in terms of electrical charges. This is accomplished by showing the number of electrons gained. Show the gain of

electrons on the left side of the equation.

Example:

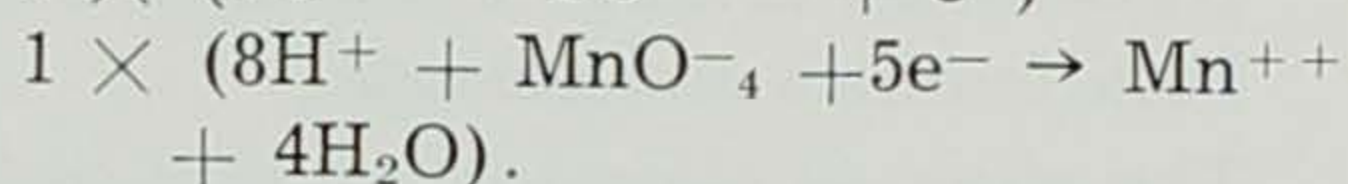
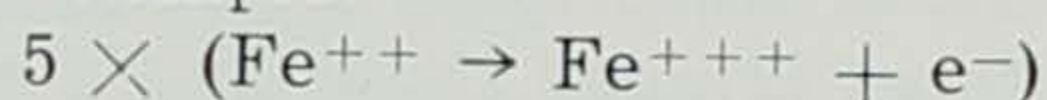


The partial equation is now balanced atomically and electrically.

e. Step 5.

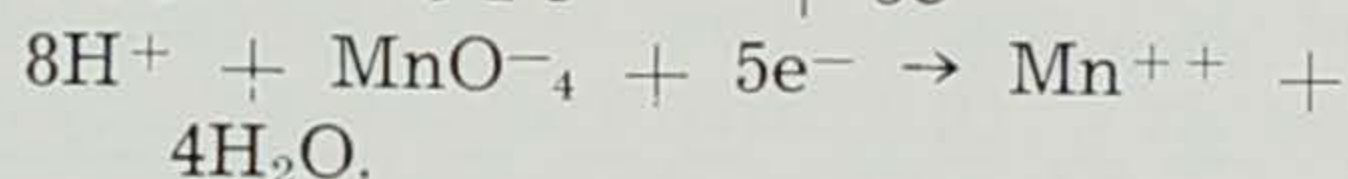
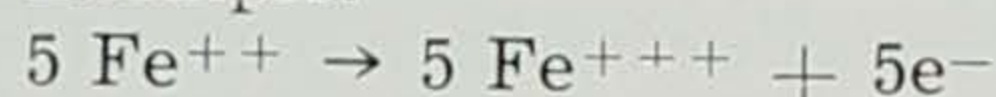
- (1) Multiply each partial equation *completely* by the number of electrons gained or lost in the other partial equation.

Example:



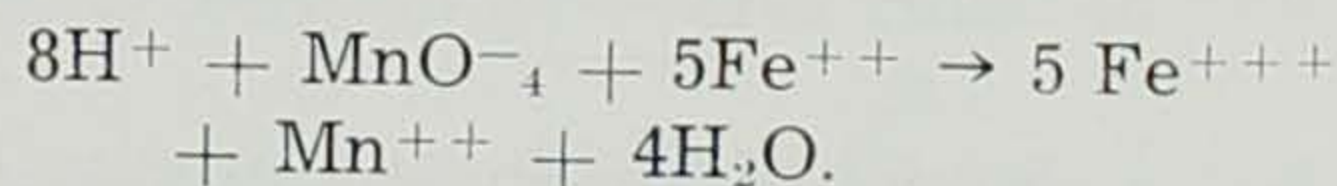
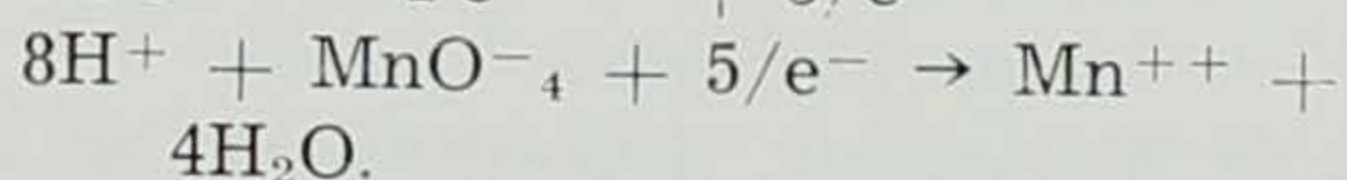
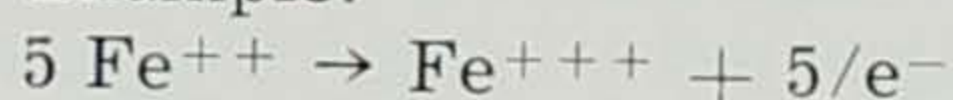
This step will result in two partial equations in which the loss of electrons in one equation is equal to the gain of electrons in the other equation.

Example:



- (2) Add the two balanced partial equations resulting from Step 4 to give the completely balanced ionic equation. Cancel electrons before adding. Cancel water molecules, H^+ AND OH^- ions, when they appear on both sides of the equation.

Example:



The ionic equation is now balanced atomically and electrically.

NOTE: By this time this unit has been covered, the pupil should have acquired skill in balancing oxidation-reduction equations of reasonable difficulty. Considerable class time should be devoted to practice under supervision and additional assignments for homework should be made. Clear understanding of oxidation-reduction reactions is necessary for

the work in electrochemistry which follows.

UNIT VIII

THE PERIODIC TABLE

A. The basic pattern.

Give a group of pupils a large number of specimens of various objects which can be arranged with respect to at least two different properties. Examples are various rocks, various woods, various colored glass objects, various sized metal washers of iron, brass, fiber, etc. Have them arrange these into a chart having both families and periods.

The effort at classification will serve to clarify the relations of the periodic chart of the elements.

As a project, class groups can make chemical periodic charts, attaching specimens of the elements or their compounds where they are available and pictures gleaned from old chemical magazines or books in place of less available elements.

1. Early attempts to classify elements.
 - a. Döbereiner, 1817—groups of three called triads.
 - b. Newlands, 1863—law of octaves.
 - c. Lothar Meyer, 1869—arrangement by atomic weights in periods.
 - d. Mendeléeff, 1869—first periodic table.
2. The periodic law. The chemical properties of elements are periodic properties of their atomic numbers.
 - a. Moseley, 1913—discovery of atomic numbers. See UNIT II.A.5.
 - b. Bohr, 1913—atomic theory. See UNIT II.A.6.
 - c. Stoner-Main Smith, 1924—table of electron arrangements of the elements.
3. The modern periodic table.
 - a. Each element assigned a definite place in the table depending upon its atomic structure.
 - b. Elements arranged into horizontal groups called periods with progressive changes in chemical properties within each period but with recurrent properties from period to period.
 - c. Elements arranged into vertical groups called families with similar properties.

- d. Different tables have different forms and arrangements of the two kinds of groups.

B. Divisions of the table—families, periods.¹

1. Periods.

Point out on the periodic chart the various properties allied to its arrangement such as electronegativities, oxidation numbers, shells of electrons, ion sizes, etc.

- a. One very short period containing only H and He corresponding to the first quantum level or shell of electrons.
- b. Two succeeding short periods of eight elements each corresponding to the second and third principle quantum levels or shells of electrons.
 - (1) From Li to Ne, the second shell of eight electrons is progressively filled.
 - (2) From Na to Ar, the third shell of eight electrons is filled.
- c. Two succeeding long periods followed later by a third, called transition series, corresponding to filling the fourth, fifth, and sixth principle quantum levels with eight electrons with expansion of the next inner quantum level to eighteen electrons.
 - (1) From K to Kr.
 - (2) From Rb to Xe.
 - (3) From Cs to Rn.
- d. Two inner transition series in which the number of electrons in the second inner quantum level or shell is progressively increased from eighteen to thirty-two.²
 - (1) Lanthanide series from elements 57-71.
 - (2) Actinide series from elements 89-102.
- e. Metallic character of elements decreases from left to right across each period.

Have pupils plot the values of one or more of the above periodic properties as ordi-

¹ W. R. Walker and G. C. Curthoys, "A New Periodic Table Based on the Energy Sequence of Atomic Orbitals," *Journal of Chemical Education*, XXXIII (1956), 69; D. DeVault, "Method of Teaching the Electronic Structure of the Atom," *Journal of Chemical Education*, XXI (November-December, 1944), 526, 575; J. A. Campbell, "Atomic Size and the Periodic Table," *Journal of Chemical Education*, XXIII (1946), 525.

² L. A. Hiller and R. H. Herber, *Principles of Chemistry* (New York: McGraw-Hill Book Company, 1960), 38-44; 30-32; Michell J. Sienko and Robert A. Plane, *Chemistry* (second edition; New York: McGraw-Hill Book Company, 1961), 46-73. Note: College level, especially good for instructor reference in this whole area of atomic theory.

nates against atomic numbers of the elements as abscissas to emphasize the periodic nature of the properties.

-
- (1) Elements at the left are more metallic and more strongly base forming.
 - (2) Elements in the center tend to be amphoteric.
 - (3) Elements at the right are nonmetallic and acid forming.

2. Families.

Have a pupil prepare to give the class a detailed résumé of the properties of a family of the elements.

Allow members of the class to add any properties missed or to criticize any erroneous additions.

If money is available for the purchase of various sizes of plastic foam balls such as are used for Christmas ornaments, have class groups prepare graphic charts of different families of the elements so that the various families, when combined, make a complete graphic chart.¹

Other material and directions for producing models of this kind are obtainable from Dr. R. T. Sanderson, Department of Chemistry, Arizona State College, Tempe, Arizona.

-
- a. Families are groups of elements with identical (or nearly identical) electron arrangements in their outermost quantum levels or shells.
 - (1) Elements in a family resemble each other closely in possible oxidation states.
 - (2) Elements in a family resemble each other closely in chemical properties.
 - (3) Elements in a family differ progressively in physical properties, atomic radius, and activity due to increasing number of filled shells as the atomic number increases.
 - b. At the left of the table, the elements of a family increase in chemical activity from top to bottom. Example: Cs is more active than Li; Ba is more active than Ca.
 - c. At the right of the periodic table, the elements of a family decrease in chemical

activity from top to bottom. Example: I is less active than Cl; both are much less active than F.

3. Value of the periodic table.

- a. A systematic organization of chemical elements with a view to their structure and properties—helps to see chemistry as a whole.
- b. An aid to the memory.
- c. (Historical) Useful for checking experimental work on physical constants.
- d. (Historical) Used to predict the presence of elements and their properties not known at the time.

UNIT IX SOME COMMON ELEMENTS AND THEIR COMPOUNDS

A. Halogens, F₂, Cl₂, Br₂, I₂.

Prepare chlorine for the class by oxidation of HCl with MnO₂. Collect several bottles of the gas, a test tube of water solution, finally leading the gas escaping from the end over CaO in the apparatus shown in FIGURE 11. This arrangement avoids escape of gas into the room. Show the reaction of Cl₂ gas with H₂ (if a cylinder is available). The jet of burning hydrogen will continue to burn in chlorine. Show reaction with sodium metal and antimony metal powder.

Show the bleaching action of chlorine water on colored cloth. (Try before class because many modern dyes are "fast" to oxidation.)

Show the properties of "chlorinated lime" produced in the end tube as a bleach and source of available chlorine. In place of CaO, one may use a solution of NaOH and make NaOCl. Relate this to commercial bleaches such as "Chlorox."

Using chlorine water, displace Br and I from solutions of their salts. Place a small amount of carbon tetrachloride in the bottom of the test tube to pick up the displaced halogen.²

Film: *The Halogens*. S.U.I. No. U 2367.

1. General characteristics of the family and oxidation states, -1, 0, 1, 3, 5, 7, common states.

¹ R. T. Sanderson, "Models for Demonstrating Electronegativity and Partial Charge," *Journal of Chemical Education*, XXXVI (October, 1959), 507-512.

² Manufacturing Chemist's Association, Inc., "Concentration of Sodium Hypochlorite Solution—Experiment No. 15," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 29.

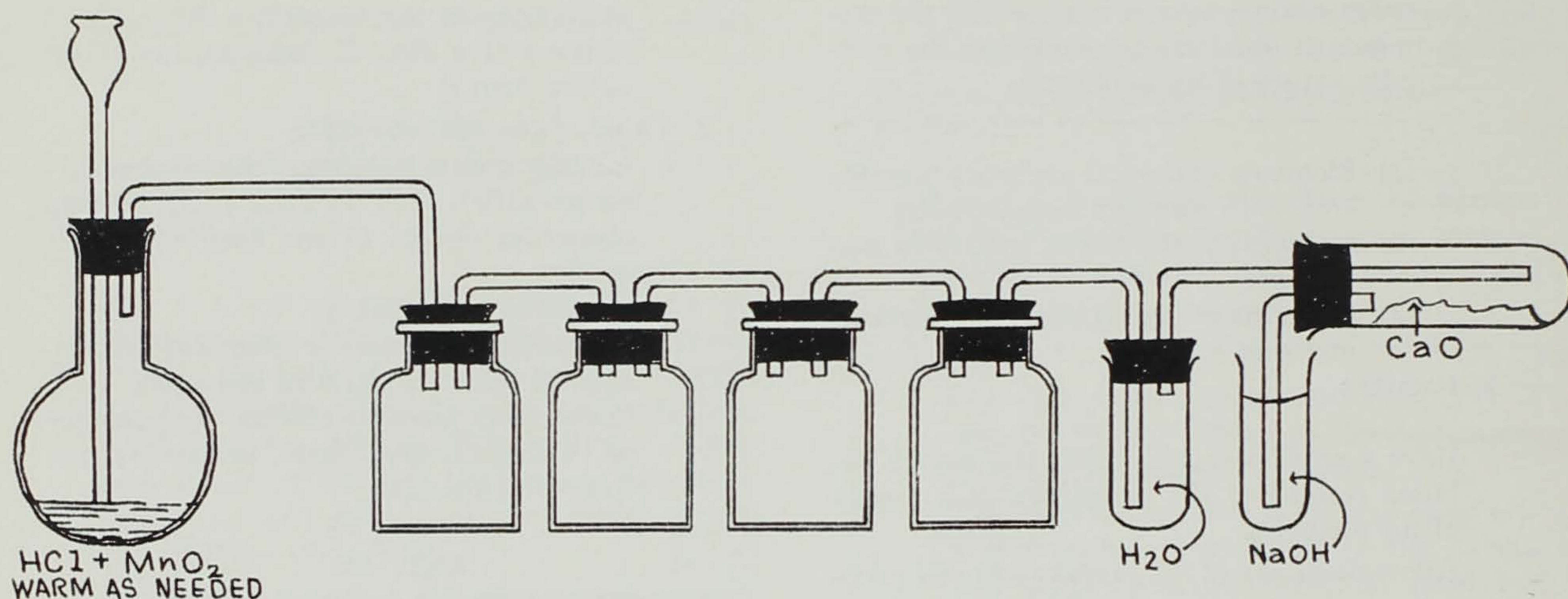


Figure 11

2. Activity series of the halogens.
 - a. Displacement series from compounds where halogen is in -1 oxidation state.
 - b. Displacement series from compounds where halogen is in positive oxidation state.
3. Fluorine—discovery, preparation, properties, uses.
4. Chlorine.
 - a. Discovery.
 - b. Sources and preparation.
 - (1) Laboratory preparation.
 - (2) Commercial production.
 - c. Properties.
 - (1) Activity.
 - (2) Electron shell structure.
 - (3) Common and special oxidation states.
 - (4) Physical properties.
 - (5) Chemical properties.
 - d. Uses—industrial chemical of major importance.
 - (1) Chlorinating agent for organic chemicals.
 - (2) Oxidizing agent.
 - (3) Bleaching and disinfecting agent.
 - (4) Production of metal halides.
 - e. Important compounds of chlorine.
NOTE: The extent to which these are studied in detail is left to the discretion of the instructor.
5. Bromine—discovery, preparation, properties, uses, important compounds.
6. Iodine—discovery, preparation, properties, uses, important compounds.

See Note: D. 3. under Nitrogen regarding NI_3 , page 38.

Pupils may be allowed to prepare both bromine and iodine in the laboratory if proper supervision is maintained.

Directions to be found in most laboratory manuals should be carefully followed and pupils should be cautioned about careful handling of products.

B. Sulfur and its family, S, Se, Te, Po, (O).

Most laboratory manuals give useful experiments on sulfur, hydrogen sulfide, and sulfate compounds.

NOTE: Hydrogen sulfide is not only offensive but very poisonous. If used from a cylinder in more than minimum quantities, take all precautions to secure excellent ventilation.¹

Film: *Sulfuric Acid*. I.S.U. No. NS-291, 16 min., b & w.

Film: *Sulfur*. 22 min., U.S. Bureau of Mines No. 188.

Film: *The Magic of Sulfur*. Texas Gulf Sulfur Co., U.S. Bureau of Mines.

Film: *Sulfur and Its Compounds*. I.S.U. NS-600, 10 min., color.

1. General characteristics of the family and common oxidation states, $-2, 0, 4, 6$.

¹ Manufacturing Chemist's Association, Inc., "Reducing Action of Hydrogen Sulfide—Experiment No. 23," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 45.

2. Sulfur.¹

Use the various compounds of sulfur to bring out the necessity of being able to infer oxidation number. Relate the electronegativity of the elements with which it combines to the role of sulfur in its combination.

a. Sources and methods of production.

- (1) Mining by conventional methods.
- (2) Frasch process mining.
- (3) Recovery as by-product of metal refining, etc.

b. Properties.

Discuss in detail the physical properties of sulfur that make it possible to use the Frasch process in mining it.

Draw sketches of the resonance structure of the oxides of sulfur. Discuss resonance. Bring out the nature of a polymer. Suggest how sulfur and its bonding structure give rise to the polymer.

Produce the various allotropic forms of sulfur. Have the pupils pay particular attention to the conditions necessary for the formation of the various allotropes.

- (1) Physical—form, allotropy, etc.
- (2) Chemical.

c. Uses.

d. Important compounds.

Use hydrogen sulfide to indicate how it produces insoluble precipitates and note how their color can be associated with identification of ions. Note here that a more modern approach is the use of thioacetamide.

Demonstrate the correct method for diluting acids. Suggest safety precautions to be taken, especially with sulfuric acid. Find empirical formula of a sulfide such as copper sulfide. Heat copper coil in excess of sulfur. Note oxidizing action of sulfur here. See this outline UNIT III.A. 2.a.

- (1) Hydrogen sulfide—properties and uses.

- (2) Sulfur dioxide and sulfurous acid—properties and uses. (See texts for expansion.)

- (3) Sulfur trioxide and sulfuric acid.²

- (a) Properties and uses of sulfuric acid.
- (b) Methods of manufacture.
- (c) Sulfate salts of importance.

3. Selenium and tellurium—brief mention of properties and uses.

C. Carbon and its family, C, Si, (Ge)—bonding, properties, and inorganic chemistry. For organic compounds see UNIT XII.

Make a drawing of the orbital arrangement of carbon. Suggest reasons for its different oxidation numbers. Mention the ground state and excited state at this point.

1. General characteristics of the carbon family.
 - a. Carbon present in all living things.

Show the presence of carbon in all living matter by immersing several organic substances in sulfuric acid.

Place 100 g of sugar in an evaporating dish. Moisten thoroughly with concentrated H_2SO_4 . This demonstrates both the presence of carbon in the sugar and the dehydrating ability of sulfuric acid.

Film: *Carbon and Its Compounds*. I.S.U. No. NS-3390, 11 min., color.

- b. Silicon, a common constituent of many rocks and minerals.
 - c. Special bonding characteristics due to peculiar electron structure.
2. Properties and uses of free carbon.

Prepare charcoal and coke. Use boneblack to illustrate the property of absorption. Note the insolubility of carbon. Show its reducing action on oxides.

- a. Allotropy—graphite, diamond (amorphous carbon); properties of various forms.
- b. Sources and preparation.
- c. Uses of carbon in various free states.

¹ Hubert N. Alyea, "Tested Demonstrations in General Chemistry," *Journal of Chemical Education*, XXXVII (1960), A298, No.s 19-165; 19-18.

² Manufacturing Chemist's Association, Inc., "Sulfur to Sulfuric Acid—Experiment No. 25," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 49.

- (1) Diamond—jewelry, cutting agent in grinding stones, phonograph needles, dies for fine wire drawing.
- (2) Graphite—for pencils, for lubricant, for brushes and electrodes.
- (3) Amorphous carbon—arc electroreducing agent in metallurgy, paint, rubber filler, absorbing and purifying agent.

3. Carbon monoxide—properties (poison).

Prepare carbon monoxide by the use of formic acid. Be careful of fumes. Note action of sulfuric acid in this reaction.

4. Carbon dioxide and carbonates.

A field trip to the local water plant, if it employs lime-carbonate softening, will illustrate uses of carbonates as well as use of gaseous CO_2 in recarbonation.

Use a compressed carbon dioxide fire extinguisher to produce a small amount of carbon dioxide snow and show how the carbon dioxide sublimates. Explain that the extinguisher works both by cooling and smothering the fire.

Show the test for carbon dioxide by the use of limewater.

Pour carbon dioxide from a beaker over a trough which contains several candles at different heights.

Have pupils find the weight of a gram molecular volume of carbon dioxide. Prepare the carbon dioxide by the action of an acid on a carbonate. Pay close attention to quantitative techniques here.

a. Method of preparation.

b. Uses.

- (1) Fire extinguishers.
- (2) Water and soft drinks.
- (3) Leavening agent in cooking.
- (4) Other uses at discretion of instructor.

See common texts.

5. Silicon, silicates, silicones. Expand if time allows.

D. Nitrogen and its family, N, P, (As, Sb, Bi).

1. General characteristics and oxidation states, —3, 0, 3, 5.
2. Sources, properties, and uses of nitrogen.
3. Important compounds of nitrogen.

Preparation and explosion of small amounts of nitrogen triiodide is interesting and spectacular, but it is our opinion that the less said about this compound the better. It is easily prepared in a clandestine manner and leads to dangerous situations in and out of class. After all, it does not add much to the chemical understanding of the elements involved.

Have pupils prepare ammonia by the action of calcium hydroxide and ammonium chloride. Have pupils note physical properties. Note that it is an anhydride.

Allow pupils to prepare NO and NO_2 following regular safe directions in any one of several manuals.¹ CAUTION THEM TO AVOID FUMES OF NO_2 AS DANGEROUS POISONS.

a. Ammonia.

- (1) Laboratory preparation.
- (2) Commercial production.
 - (a) Haber process.
 - (b) Cyanamide process (now nearly obsolete).
- (3) Properties and uses.

b. Nitrogen oxides.

Nitrous oxide, N_2O , may be prepared by very gently heating pure NH_4NO_3 in a test tube generator. The heating should be done very gently. CARE. *Rapid heating may cause an explosion!* DO NOT UNDER ANY CIRCUMSTANCES ALLOW ANY OF THE MOLTEN MATERIAL OR FOAM FROM IT TO TOUCH THE RUBBER STOPPER. Collect the gas over water. Pupils should note the color and may note the odor. Plunge a glowing splinter into the gas and observe that the splinter is relighted in the same way as with oxygen, showing the ready tendency of this gas to dissociate into N_2 and O_2 . Demonstrate pressurized can of whipped cream in which the carrier gas is N_2O . Nitric oxide, NO , and nitrogen dioxide, NO_2 , may also be prepared.

¹ Harper W. Franz and Lloyd E. Malm, *Essentials of Chemistry in the Laboratory* (San Francisco: W. H. Freeman and Company, 1961). Experiment 23 is good. See also L. E. Young and W. M. Petty, *Chemistry for Progress Laboratory Manual* (Englewood Cliffs, New Jersey: Prentice-Hall, Inc., 1959). Experiments 38 and 39.

Film: *Ammonia*. I.S.U. NS-262, 20 min., b & w.

Film: *Nitric Acid Compounds and the Nitrogen Cycle*. I.S.U. NS-559, 18 min., b & w.

Laboratory preparation and properties.

c. Nitric acid.

Prepare concentrated HNO_3 by distillation of a small amount into a water cooled test tube from a small distilling flask containing 5 g dried NaNO_3 well moistened with concentrated H_2SO_4 . Note some decomposition. Hold the test tube of pure HNO_3 distillate in a holder and boil the acid, meanwhile holding a tuft of wool over the end of the tube in the fumes. The wool will first turn yellow (xanthoproteic acid reaction), then burst into flames, showing intense oxidizing power of the HNO_3 vapor.

Test reaction of concentrated HNO_3 on a less active metal as Cu and on a more active metal as Zn or Mg. In the latter case, test escaping gas for ammonia with red litmus or more sensitively by a wad of cotton soaked in a mixture of AgNO_3 solution and tannic acid solution. NH_3 turns this mixture brown and may show silver, mirror-like surface stain.¹

- (1) Laboratory preparation.
- (2) Commercial production and uses.
- (3) Nitrates.

4. Important compounds of phosphorus.

Phosphorus and phosphates. These experiments can be done individually or by small groups of pupils.

Place a piece of red phosphorus the size of a match head in a deflagrating spoon or in the center of a tight spiral bent at the end of an 18-inch length of iron wire. Light this and lower it into a 6-ounce, wide-mouth bottle. Partly cover the bottle with a piece of glass. What are the white fumes? Write the equation. When the reaction is complete, remove the wire and burn off any remaining phosphorus in a Bunsen burner flame. Add 5 ml distilled water to the bottle and shake until

fumes are dissolved. Test the solution with litmus paper. What has been formed? Take 1 ml of the solution in a test tube and add AgNO_3 solution. Dissolve 0.1 g Na_3PO_4 or other known orthophosphate in 1 ml of water and add AgNO_3 solution. Compare result with that from the solution of the white fumes. What was formed when white fumes dissolved in water? Write equation. Test various fertilizers with AgNO_3 to determine whether they contain orthophosphates.

a. Phosphorus sesquisulfide—matches.

Demonstrate differences between white and red phosphorus. Dissolve a small ($\frac{1}{2}$ g) bit of white phosphorus in 2 ml CS_2 and pour on two sheets of filter paper supported on a tripod. It will burn spontaneously after solvent has evaporated. Repeat with the same quantity of red phosphorus. Show difference in heads of safety and "strike anywhere" matches. Where is the phosphorus in each?

b. Oxides and acids.

c. Phosphates—fertilizers.

Discuss great value of phosphate compounds as fertilizers. What are the sources of phosphorus in the United States; in the world?

E. The active metals, alkalis, and alkaline earth families.

1. General characteristics of the alkalis, (Li) Na, K, (Rb, Cs).
2. Occurrence of Na and K as salts.
3. Preparation and properties of Na and K.
4. Uses of Na and K.
5. Important compounds of Na and K.
6. General characteristics of the alkaline earths Ca (Sr, Ba).
7. Preparation and properties of Ca.
8. Important compounds of Ca.

F. Some light metals—magnesium, aluminum, zinc.

Obtain samples of Mg, Al, and Zn. (Mg samples may be obtained from Dow Chemical Company, Midland, Michigan. Other samples may be obtained locally.) Compare their properties. To

¹ Manufacturing Chemist's Association, Inc., "Preparing Nitric Acid and Finding Its Concentration—Experiment No. 26," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 51.

show the relative activity of these, clean the surface of each specimen by scraping and amalgamate by rubbing with mercury. This avoids the effect of the adherent coating of oxide on the Al and Mg. Note rapid oxidation of both. Burn Mg wire in air. Will Al wire do the same? Why not?¹

Nearly every laboratory manual for high school use has suitable experiments dealing with properties of the common metals and their salts. It is not considered necessary to list these specifically. In choosing experiments from the many available on the common metals, three criteria should be considered: (a) Does the experiment illustrate an important physical or chemical property of the metal or its compound? (b) Is the experiment so designed as to make this property or the distinction between properties more understandable rather than to further confuse the pupil? (c) Is the experiment safe in the hands of pupils and well enough described so that they can follow the directions largely on their own initiative?

1. Magnesium.

Show film, "Magnesium from the Sea." With the aid of the pupils in class, work out the equations for the process. Do the same for aluminum.

- Occurrence.
- Commercial production (from sea water), properties, and uses of the metal.
- Useful salts and their properties.

2. Aluminum.

Recall the story of the Hall process for reduction of Al_2O_3 ore to Al. Perform the thermite reaction.²

- Occurrence and important ores.
- Commercial production, properties, and uses.
- Useful salts and their properties.

3. Zinc.

Discuss reduction of ZnO ore with carbon—

¹ Manufacturing Chemist's Association, Inc., "Chemical Properties of Metals—Experiment No. 27," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 53.

² See Dull and others, *Modern Chemistry* for this.

a typical metallurgical process. This is an excellent opportunity to relate the activities of the various metals to the difficulty of reduction of their compounds. Continue this through the heavy metals which follow. Emphasize the differing energies per mole required for the various reductions.

- Occurrence and important ores.
- Metallurgy, properties, and uses.
- Useful salts.

G. Some important heavy metals—copper, iron.

See F, this unit, for experiments.

- Transition metals in general. These are elements having one or more incomplete inner shells of electrons in their structure. There are about 55 elements in this category.

Discuss the properties of the transition metals in relation to their atomic structures and their positions in the periodic table.

Discuss the position of these elements on the periodic chart and show the structures of the shells and orbitals.

Discuss the concept of the triad arrangement as found with Fe, Co, Ni.

- All transition elements are metallic in the free state.
- Transition elements show similarity in physical and chemical properties to their neighbors.
- Show variable oxidation states with two and three being common.
- Many of their compounds are colored.

Show that the compounds of these ions are colored. Use copper, cobalt, nickel, chromium, and iron.

Exercises in identification of metals and their salts.

Use the Borax bead test. Use the cobalt nitrate test for metals and flame tests. Supply several metals and have the pupils relate their position in the periodic chart to their properties and those of their common compounds.

- Form complex ions.

Discuss the formation of complex ions. Any college text should give a good basis for this discussion.

— — — — —

- f. Both covalent and ionic bonds are formed by ions of transition elements.
2. Copper.
- — — — —

The metallurgy of copper is most interesting. Outline the various steps of the process and give equations representing as simply as possible the chemical reactions taking place in (a) roasting, (b) Bessemer converting, (c) fire refining including poling, and (d) electrolytic refining.

This is a good place to point out that much high grade ore has been used. The processes for concentration and beneficiation of lower grade ores are becoming more important. Describe the flotation process for concentrating copper ores, and the beneficiation of Taconite iron ores.

— — — — —

- a. Occurrence and important ores.
b. Metallurgy, properties, and uses.
c. Useful salts.
3. Iron.
- — — — —

See G. 2.

Discussion of hardening and of tempering. A visit to a foundry or a local blacksmith is suggested to accompany a discussion of the nature of these elements.

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- a. Occurrence and important ores.
(1) Hematite ores.
(2) Taconite ores—their beneficiation.
- b. Metallurgy and properties.
(1) Production of pig iron—the blast furnace.
(2) Production of steel.
(a) Bessemer process.
(b) Open-hearth process.
(c) Electric furnace.
(3) Alloy steels.
- c. Uses of iron and steel.
- — — — —

Discuss the Domain theory of magnetism.

— — — — —

Film: *Pioneer of Progress* (steel). I.S.U. NS-1178, 14 min., b & w.

Film: *Stainless Steel*. I.S.U. NS-5003.

Film: *The Drama of Steel*. 34 min. (1946), United States Bureau of Mines No. 218.

d. Useful salts of iron.

UNIT X

EQUILIBRIUM AND KINETICS

Equilibrium is a steady state in which opposing forces are exactly balanced. Kinetics refers to the speed with which equilibrium is approached.

A. Principles governing equilibrium.

1. Types of equilibrium.

a. Physical.

- (1) Static—Example: A funnel sitting on its top.
(2) Dynamic—Example: A teeter-totter when in balance.
- — — — —

Discuss various physical equilibria and demonstrate with a funnel, teeter-totter, etc. A 250 ml round bottom flask with a layer of shot in the bottom held in place with paraffin is useful.

— — — — —

b. Chemical.¹

— — — — —

Stress chemical equilibrium as a dynamic condition. Have pupils illustrate your point in various ways. For example, have one pupil place the palms of his hands against those of another and push as the other pushes back, thus illustrating a steady state of equilibrium with the forces still in action but in opposite directions. Again, the point may be illustrated as follows: Place two large beakers three-fourths full of water on the lecture room table. Have one pupil ladle water from the first beaker to the second and another ladle water in the opposite direction. At first have them use identical ladles and work at the same speed. Small kitchen measuring ladles in one-half cup size work very well. Under these conditions the volume in the two beakers will remain nearly the same. Next have them illustrate the effect of physical factors on the equilibrium by having one use a one-fourth cup or a three-fourths cup ladle but still maintain the same speed of ladling. Call attention to the fact that at first

¹ W. R. Carmody, "Dynamic Equilibrium," *Journal of Chemical Education*, XXXVII (July, 1960), 312.

there results a change in relative volumes in the beakers, but soon a steady state again ensues for as the water level in one beaker becomes low, the pupil with the larger ladle cannot completely fill it with each transfer. Different speeds can be tried. Other variations are possible.

(1) Dynamic only.

(a) Homogeneous—all in same state of matter. Example: $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$ all gases.

(b) Heterogeneous—between different states of matter or phases. Example: $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$.

(2) A chemical equilibrium is the state of balance attained when the speed of the forward reaction is exactly equal to the speed of the backward reaction.

2. Speed of attainment of equilibrium depends upon the speeds of the reactions involved.¹

Film: *Velocity of Chemical Reactions*. I.S.U. NS-1096, 11 min., b & w. Also, S.U.I. No. U-782, 11 min., b & w. Sound.

a. Speed of reactions depends upon the frequency of collision of the reacting molecules.

The molecular model suggested in the supplementary lecture on "The Behavior of Gas Molecules" may be used at this point to illustrate the fact that the speed of reactions depends upon the frequency of molecular collision. Stress that only those molecules having kinetic energies higher than a certain threshold value (energy of activation) have inelastic collisions resulting in chemical reaction. Raising the temperature not only increases the average kinetic energy but also increases the fraction of molecules having the higher kinetic energies. This accounts for rise in temperature speeding up reactions.

¹ Manufacturing Chemist's Association, Inc., "Rates of Reaction—Experiment No. 11; Concentration and Decomposition Rate—Experiment No. 17; Effect of Temperature on Rate of Reaction—Experiment No. 18; Effect of Concentration on Rate of Reaction—Experiment No. 16," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 21, 33, 35, 31.

b. The collision efficiency, i.e., the fraction of collisions which result in chemical change or reaction.
c. Physical factors affecting the speed of a reaction.²

Class project or group experiment. Purpose: To prove that a positive catalyst speeds up a reaction without affecting the quantity of product formed. This is to be preferred over the Manufacturing Chemist's Association, Inc., alternate. Use the apparatus shown below (FIGURE 12).

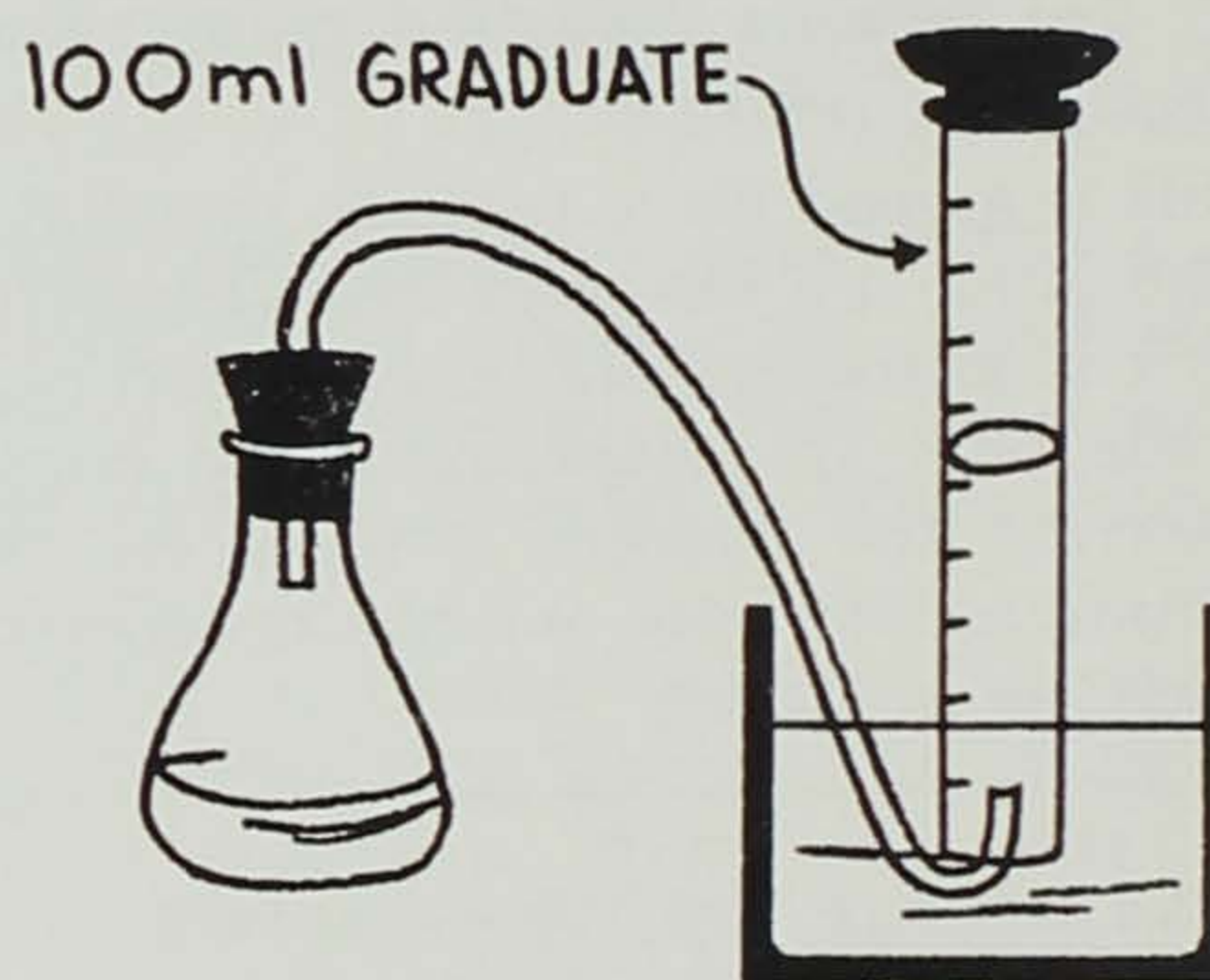


Figure 12

Give to each pupil doing the experiment the same volume (about 8 ml) of 3 per cent H_2O_2 but a different very small weighed amount of MnO_2 catalyst. The instructor will have to determine this weight of the smallest amount to be used as MnO_2 varies in catalytic activity. This amount should just give a slow but steady evolution of O_2 . Have the pupil initiating the experiment first add his H_2O_2 to the flask, then 10 ml of water, and finally the MnO_2 . The flask should be stoppered at once and the time observed. The flask should then be shaken gently and continuously until no more gas is evolved. The pupil with the next larger weight of MnO_2 should then repeat the experiment, etc. Have the class record the data from each experiment. The time to reach a certain predetermined volume

² Julian R. Brandon, "Introduction to Le Châtelier's Principle," *Journal of Chemical Education*, XXXVIII (1961), A491; Manufacturing Chemist's Association, Inc., "A Study of Catalysis—Experiment No. 19; Effect of Weight of Catalyst on Reaction Rate—Experiment No. 20," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 37, 39; Richard B. Escue, "Secular Equilibrium," *Journal of Chemical Education*, XXXVII (October, 1960), A677.

should be observed, and the final volume reached at "infinite" time should also be recorded. These final volumes should all be approximately the same. The weights of MnO_2 used and the times to reach the predetermined volume from all the experiments should be made into a table. Each class member should then plot weight of MnO_2 vs. time, weight of MnO_2 vs. $1/\text{time}$, and weight of MnO_2 vs. $\log \text{time}$.

- (1) Nature of the reactants.
- (2) The amount of surface.
- (3) The effect of temperature—rise in temperature increases speed.

Film: *Catalysis*. S.U.I. No. U-785, 11 min., b & w. Sound.

- (4) Effect of concentration.
- (5) Use of a catalyst—a catalyst is a substance which affects the rate of attainment of equilibrium without affecting the end result.

3. Behavior of systems in equilibrium.

- a. Described by the general principle of Le Châtelier—WHEN A DISTURBING INFLUENCE (temperature, pressure, or concentration) AFFECTS A SYSTEM THAT HAS ATTAINED EQUILIBRIUM, THE SYSTEM WILL REACT IN SUCH A WAY AS TO DIMINISH THE EFFECT OF THE INFLUENCE.

- b. Disturbing influences illustrated.
-

Illustrate the application of factors that affect the rate of reaction by describing the synthesis of ammonia. Create class discussion by asking questions concerning the factors involved.

Continue the discussion to the point where the pupils understand the effect of temperature, pressure, and concentration upon the yield of this reaction.

- (1) Pressure—example: $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$. Here four volumes of gases react to form only two volumes of ammonia gas. Increased pressure will tend to force the reaction toward completion since the decrease in volume will diminish the effect of the increased pressure.

- (2) Temperature—example: $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$ $\Delta H = -10.9 \text{ kcal}$.

(NOTE: Minus sign indicates that heat is given off when the reaction goes from left to right.) A rise in temperature will tend to force this reaction toward the *left* with a *decrease* in amount of products since such a backward reaction will absorb heat diminishing the effect of the rise in temperature.

- (3) Concentration.

- (a) Effect of small changes in concentration is to cause small changes in extent of the reaction. In the above example, increasing the concentration of hydrogen will cause the formation of more product since this will use up hydrogen and thus diminish the effect of the increased concentration.

- (b) Effect of large changes, i.e., virtual removal of product. Example: Removal of ammonia as fast as it is formed will cause the reaction to go nearly to completion. NOTE: Products may be removed as *gases*, as *precipitates*, or sequestered as *little ionized substances* (molecules) or as *chelate complexes*. All these factors tend to cause nearly complete reactions.

B. How chemists use factors affecting equilibrium to gain their ends.

In DeBruyne, Kirk, and Beers, *Semimicro Chemistry*, "Chemical Equilibrium," p. 133, No. 31, part one shows the effect of various concentrations of FeCl_3 and NH_4SCN on the color of FeSCN^{++} .

Part two shows the effect of the concentration of HCl on the solubility of NaCl .

1. Qualitative applications.

Discuss the law of chemical equilibrium qualitatively showing the effect on the concentrations of the other ingredients; of varying the concentration of one of them. This can be done on the board by changing the

size of the formulas used in the mathematical expression.

- a. Gases—example: In the Haber process for the production of ammonia, high pressures are used to secure greater yields of ammonia. See UNIT X.A.3.b. (1).

$$\frac{[\text{NH}_3]^2}{[\text{N}_2] [\text{H}_2]^3} = K_{\text{equil.}}$$

- b. Molecular solutions—example: $\text{C}_2\text{H}_5\text{OH} + \text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5 + \text{H}_2\text{O}$. Increase in concentration of ethyl alcohol ($\text{C}_2\text{H}_5\text{OH}$) will cause greater formation of the product, the ester, and methyl acetate.

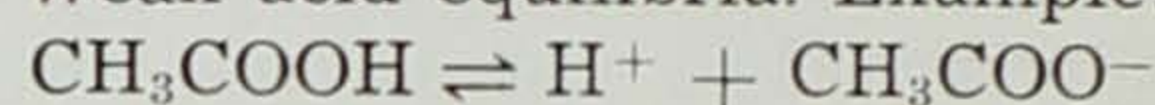
$$\frac{[\text{CH}_3\text{COOC}_2\text{H}_5] [\text{H}_2\text{O}]}{[\text{C}_2\text{H}_5\text{OH}] [\text{CH}_3\text{COOH}]} = K_{\text{equil.}}$$

- c. Ionic solutions. SPECIAL IMPORTANCE.

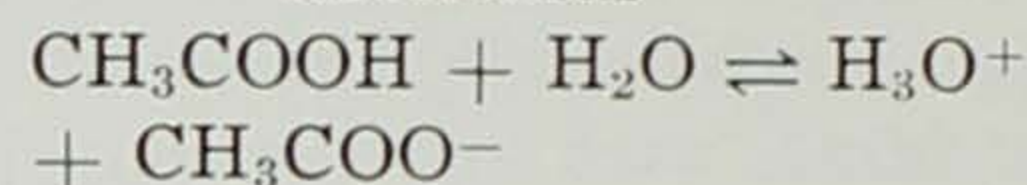
Illustrate the common ion effect by the use of calcium carbonate and acetic acid. To a beaker which contains calcium carbonate and a small amount of water, add 2 N acetic acid. To a similar beaker containing approximately the same amount of calcium carbonate and water add 2 N acetic acid in which there has been dissolved an equivalent amount of sodium acetate. Both additions will cause the carbonate to foam, but the one which has the pure acetic acid in it will foam most rapidly. The sodium acetate reduces the availability of the hydrogen ions. Equations for the reaction with the carbonate and for the effect of the sodium acetate upon the acetic acid should be shown. Note that the final amount of foam in each reaction will be the same since the total number of hydrogen ions available in each case is the same.

If a suitable acid-base indicator of wide range, such as Hartman-Leddon Company Universal Indicator, is available, use the indicator on the 2 N acetic acid with and without the sodium acetate present to show the effect of the common acetate ion upon the pH.

- (1) Weak acid equilibria. Example:



Arrhenius

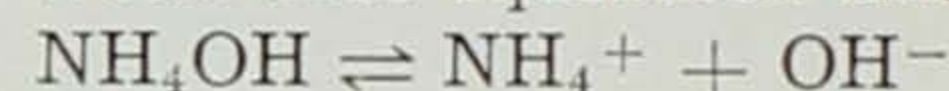


Brönsted

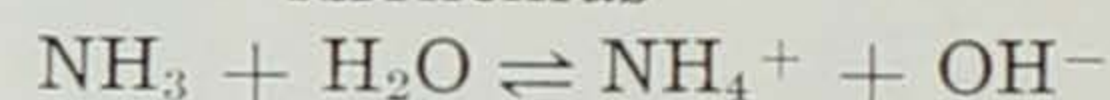
Increasing the concentration of hydrogen ions will decrease the concentration of acetate ions and increase the concentration of molecular acetic acid.

Increasing the concentration of acetate ions will decrease the concentration of hydrogen ions and increase the concentration of molecular acetic acid. This is common ion effect.

- (2) Weak base equilibria. Example:



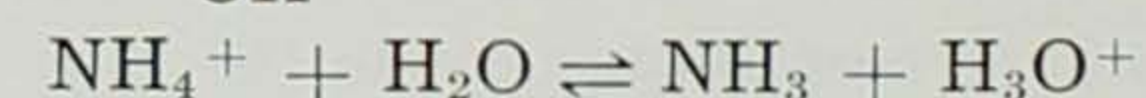
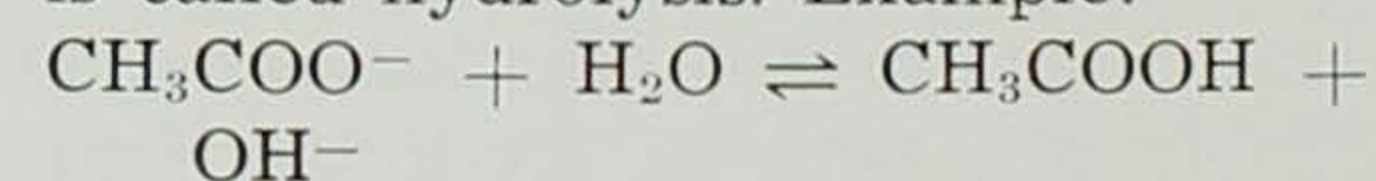
Arrhenius



Brönsted

Increasing the concentration of hydroxide ions will decrease the concentration of ammonium ions and increase the concentration of molecular ammonia. (Note the existence of ammonium hydroxide molecules is hypothetical. The postulation of this compound is not needed according to modern theory.) Increasing the concentration of ammonium ions will decrease the concentration of hydroxide ions and increase the concentration of molecular ammonia. This is common ion effect.

- (3) Hydrolytic equilibria. The neutralization of a weak acid by hydroxide ions or of a weak base by hydrogen ions is not a complete reaction but reaches equilibrium before completion. This is because the weak acid or base is less than completely ionized as is water. This results in a competition for protons according to Brönsted. Similarly, when the anions of a weak acid or the cations of a weak base are dissolved in water, the reverse of the neutralization reaction occurs. This is called hydrolysis. Example:



The result is that salts of weak acids give slightly alkaline solutions in

water while salts of weak bases give acid solutions.

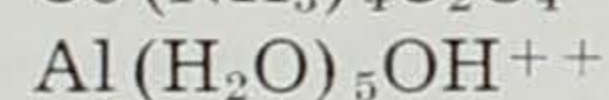
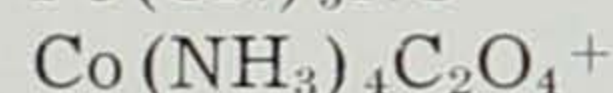
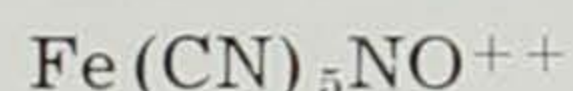
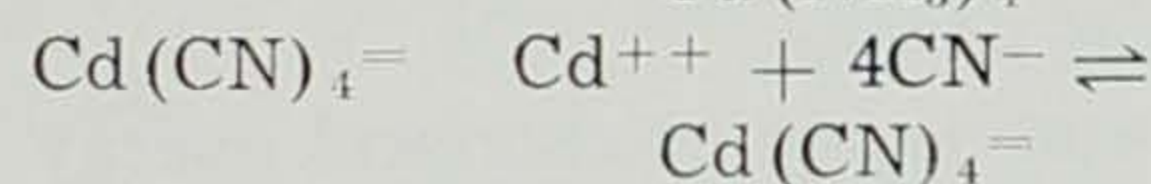
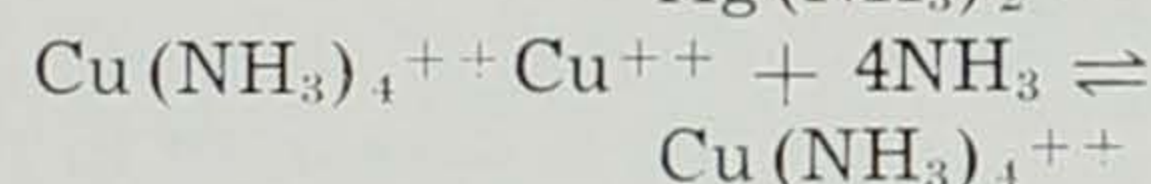
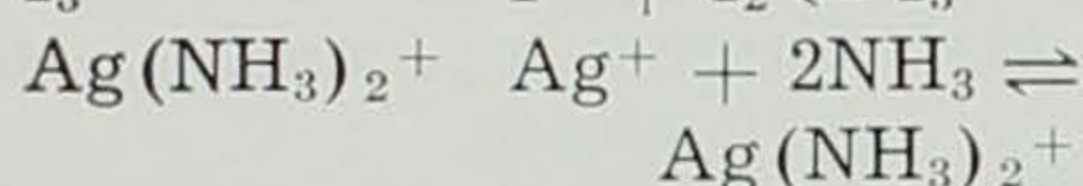
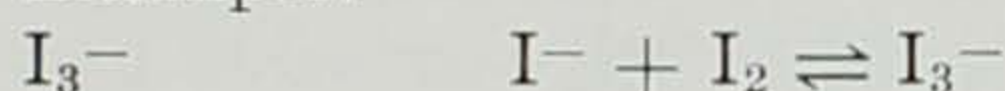
d. Heterogeneous systems—example:

$\text{AgCl(s)} \rightleftharpoons \text{Ag}^+ + \text{Cl}^-$. Increasing the concentration of chloride ions will decrease the concentration of Ag^+ ions which can remain in solution at equilibrium and decrease the solubility of AgCl solid, forming more precipitate.¹

e. Complex ion equilibria.

(1) Nature of complex ions. Complex ions are formed by the combination of ions with ions or of ions with neutral molecules. Such combinations are held together by coordinate covalences. One member of the combination must have an unused pair of electrons and the other must be an electron pair acceptor. Transition metals are very prone to form complex ions because they are either electron deficient or can expand their shells to accommodate more electrons.

Example:



(2) Structure and coordination number.

(a) Ions and molecules forming coordination compounds have definite special relations and form geometrical isomers.

(b) The capacity of a central ion to hold attached groups is called its coordination number. Six is a common number, but numbers of 1, 2, 4, 6, and 8 are known.

(3) Complex ion equilibria. Complex ions vary in stability over a wide range. Some are very stable as in Fe(CN)_6^{--} ; others are very unstable as in Fe(SCN)^{++} . The tendency of a complex ion to dissociate

can be expressed mathematically as can other equilibria. The equilibrium constants for such reactions are called instability constants.

$$\frac{[\text{Fe}^{+++}] [\text{SCN}]}{[\text{Fe(SCN)}^{++}]} = K = 3.3 \times 10^{-2}$$

2. Quantitative calculations.

NOTE: Quantitative calculations of chemical equilibria may be simple or very complex depending upon the situation. The extent to which instruction in calculations is advisable must rest with the individual instructor. What follows merely indicates possibilities.

a. Gaseous equilibrium calculations—ammonia or hydrogen iodide systems are well understood and fully documented.

— — — —

Bodenstein, in investigating the equilibrium conditions of the reaction $2\text{HI} \rightleftharpoons \text{H}_2 + \text{I}_2$, used volumes measured at 0° C and 1 atm to express the composition of the equilibrium mixture. His data are as given in the table below.

DECOMPOSITION OF HI AT 448° C

I_2 in cc	H_2 in cc	HI in cc	K
0.12	5.28	5.64	0.0199
0.55	3.19	9.49	0.0195
2.53	1.33	12.47	0.0185
6.97	0.65	14.93	0.0203
19.76	0.25	15.54	0.0205
25.40	0.19	15.40	0.0204
			0.0198(4) Av.

Several problems or some good board work can be based upon this data.

— — — — —

b. Molecular solutions. Equilibrium constant for formation of ethyl acetate, (see UNIT X. B.1.b.) is $K = 4$ at 25°.

— — — —

Hydrolysis of ethyl acetate, $\text{CH}_3\text{COOH} + \text{C}_2\text{H}_5\text{OH} \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5 + \text{H}_2\text{O}$

Starting amounts:

3 moles 3 moles 0 mole 0 mole

Equilibrium amounts:

$3 - 2 = 1$ mole 1 mole 2 moles 2 moles
Since all are in the same volume, the molar concentrations will be $1/V$, $1/V$, $2/V$, and $2/V$ where V is the total volume of all reagents.

¹ B. H. Peterson, "The Solubility Product of Copper Iodate: An Experiment," *Journal of Chemical Education*, XXXIV (December, 1957), 612.

$$\frac{[\text{CH}_3\text{COOC}_2\text{H}_5] [\text{H}_2\text{O}]}{[\text{C}_2\text{H}_5\text{OH}] [\text{CH}_3\text{COOH}]} = \frac{2/V \times 2/V}{1/V \times 1/V} = K = 4.$$

Berthelot and St. Gilles found other data for the reaction as follows:

Moles of alcohol	Moles of ester
0.05	0.05
0.18	0.171
0.25	0.226
0.50	0.414
1.00	0.665
2.00	0.858

- c. Ionic solutions. Simple calculations of ionization of weak acids such as acetic acid or of weak bases as ammonia are advised.¹
- d. Heterogeneous systems. Solubility product calculations may be done at discretion of the instructor. Extensive calculations in this area are not advised.

UNIT XI ELECTROCHEMISTRY

A. Production of electricity.

1. Production of an electric current.

a. Electricity can be produced by:

- (1) Movement of a conductor with respect to magnetic field, e.g., a generator.

The production of electricity by use of a generator can be demonstrated by using an old telephone generator. Hooking the generator to a neon bulb or oscilloscope will help demonstrate the production of electricity. Call attention to increased energy required when the generator is connected to bulb over that when it is not connected.

- (2) Conversion of heat into electricity, e.g., a thermocouple.

A thermocouple can be made by twisting to each end of a 3 foot iron wire a 3 foot length of bare copper wire. The free ends of the copper

wire are fastened to a galvanometer. One copper-iron junction is placed in an ice water bath. The other copper-iron junction is heated in the flame of a Bunsen burner.

NOTE: If possible, the instructor should demonstrate one or two of the methods. A brief explanation should accompany each demonstration.

Thermocouple Conversion of Heat to Electricity. This project appears in *The Science Experimenter*, No. 5631. The publisher of this magazine is The Science and Mechanics Publishing Company, 450 East Ohio Street, Chicago 11, Illinois. The article discusses the construction of various kinds of thermocouples. Enough theory is provided to make the project a worthwhile one. The article gives details for the construction of a thermopile that will operate a small electric motor.

- (3) Conversion of radiant energy, e.g., a solar cell.
- (4) Conversion of nuclear energy, e.g., a nuclear cell.
- (5) Conversion of chemical energy. This method is developed in this unit.

B. Electric cells.

1. Conversion of chemical energy into electrical energy.

a. Batteries—B.C.

- (1) Replica of 2,000 years old wet cell produced electricity.²

The instructor should refer to high school physics books for details on the production of electricity by various devices.

- b. Galvani and animal electricity.
- c. Volta.

Construct a simple Voltaic pile.

- (1) Volta showed that the Galvani phenomena was due to the presence of two dissimilar metals.
- (2) Volta's battery led to experimentation involving the construction of other kinds of cells.

¹James N. Butler, "An Approach to Complex Equilibrium Problems," *Journal of Chemical Education*, XXXVIII (March, 1961), 141-143.

²"Batteries," *The Laboratory*, XXV (New York: Fisher Scientific Company, 1956), pp. 112-113.

Show a dissected hearing aid battery or "layer built" B battery to show that the voltaic pile is still being used.

— — — — —

- (3) Volta's battery was used to produce chemical changes by means of an electric current.
- (4) Electrolysis of fused salts enabled Davy to discover new elements.

d. Difficulty of recognizing that chemical changes produced electrical energy.

- (1) It was apparent to the people who used voltaic piles that electrical energy would produce chemical changes.
 - (2) It was very difficult for them to recognize that chemical changes were the source of the energy of a voltaic pile.
- — — — —

Electricity by a Chemical Reaction.
This demonstration uses the electricity generated by a copper-magnesium cell to ignite a photoflash bulb.

— — — — —

Film: *The Electron, An Introduction*—1944, U.S. Office of Education. OE175 16 min., b & w. Nature of electrons, electron flow, emf., effect of magnetic fields.

Film: *Electrochemistry*. I.S.U. NS-3468, 11 min., b & w.

C. Cells.

1. Current producing cells.

a. Chemical cells (Voltaic).

(1) Active electrodes.

An example of this type is the dry cell which has one active electrode.

(2) Inert electrodes.

An example of this type of cell is the dry cell, one electrode of the dry cell is inert.

(3) Concentration cells.

(a) It is doubtful that this cell should form a part of this unit. The instructor should have knowledge of this cell.

(b) See reference for concentration cell discussion.

2. Electrolytic cells.

- a. Electricity is used to produce a chemical change. Electrolytic cells will be discussed under electrolysis.

D. Theory and design of voltaic cells.

1. Theory.

a. Metal vs. ion equilibrium.

— — — — —

Construction of a simple cell. Set up a 600 ml beaker containing a solution of 1 M zinc sulfate. Place a large zinc electrode in the solution. Diagram the half-cell on the blackboard. Discuss the metal ion equilibrium taking place. Commercial devices are now available to supplement the blackboard diagram. Magnetic discs are used to show transfer of electrons and ions. See also the *Journal of Chemical Education* for a similar device.

— — — — —

- (1) Discuss the reactions taking place when a zinc electrode is immersed in a solution containing zinc ions.
 - (2) Discuss the reactions occurring when a copper electrode is immersed in a solution containing copper ions.
- — — — —

Repeat above demonstration using 1 M cupric solution and copper electrode.

— — — — —

- (3) Discuss the temporary flow of electrons when the half-cells from a and b are connected by the leads of a voltmeter.
- — — — —

Connect the two half-cells from A-1a A-1b to the leads of a high resistance voltmeter. The resistance of the voltmeter must be at least 1,000 ohms/volt.

— — — — —

- (4) Discuss how the salt bridge enables the flow of electrons outside the cell to continue by keeping the net charges in the beaker equal to zero.
- — — — —

Attach a salt bridge, KCl, to the beakers. A small flask light bulb, 1.5 v., may be substituted for the voltmeter. See F.1.a., this unit, page 49.

— — — — —

2. E.M.F. of a cell.

- a. The E.M.F. of a cell depends on what reactants take part in the reactions at the electrode.

— — — — —
The pupil should be provided with zinc, copper, aluminum and tin strips, solutions of 0.5 M zinc sulfate, copper sulfate, tin chloride, and aluminum chloride. Salt bridges filled with potassium chloride can be made by the pupil. The pupil should construct and measure the E.M.F. of all possible cells. Diagrams should be made of each half-cell.

- — — — —
b. The E.M.F. of a cell is also determined by the concentration of the reactants.

— — — — —
Effect of a change in concentration of the reactants on the E.M.F. of a cell. The pupil should use the copper-zinc cell for this. A solution of ammonium hydroxide is added to the beaker containing the copper ions. Water equal in volume to the ammonium hydroxide should be added to zinc half-cell. An interesting variation is to have the pupil substitute a brass electrode for the copper electrode.

- — — — —
c. The E.M.F. is independent of the size of the electrodes and size of the cell.
d. See polarization for effects that may reduce the original E.M.F. of a cell.
3. Current output of a cell.
a. Current output of a cell depends upon the internal resistance of the cell.
(1) The size of electrode determines the current output of a cell.
(2) The concentration of the reactants determines the current output of a cell.
4. Energy output of a cell.
Total energy output of a cell depends upon the amount of chemical material available.
5. Polarization.

— — — — —
If a potentiometer is available, have pupils measure the voltage of a dry cell, preferably an old one. Next, leaving the potentiometer connected, add a low resistance of large current carrying capacity, e.g., 0.3 ohm, 10 watt, across the terminals of the cell. Measure the voltage again. Why is the voltage lower in

this case? Calculate the apparent internal resistance of the cell. After five minutes, measure the voltage a third time. Is this still lower? Explain the effect of polarization on the voltage and current supplying ability of the cell.

- — — — —
a. The gas formed, usually hydrogen, in many voltaic cells increases the internal resistance of the cell.

— — — — —
Construct a copper-zinc sulfuric acid cell. Note the rapid drop in current as the cell operates. Compare behavior of this cell to that of the Daniell cell.

- — — — —
b. Hydrogen collecting on the electrode forms a hydrogen electrode. The hydrogen electrode formed will create an E.M.F. that opposes the E.M.F. of the original cell.

— — — — —
Construct a mercury-zinc very dilute hydrochloric acid cell. Polarization is so great that almost no current is produced by this cell. Addition of a small amount of mercurous chloride to the cell will cause the voltage to jump up to original value.

- — — — —
c. A heavy current drain will change the concentration of ions in the vicinity of the electrode. This will create internal concentration cells that will affect the E.M.F. of the original cell.

— — — — —
An interesting variation in the copper-zinc sulfuric acid cell is to oxidize the copper electrode in the Bunsen flame. Compare the current output with that of the nonoxidized copper electrode. The copper oxide is soon destroyed and the current output drops again.

E. Types of cells.

1. Characteristics of commercial cells.
Commercial cells should be low in cost, have long life, should be reversible and portable.
2. Dry cells.
 - a. LeClanché, zinc-carbon cell. Construction of this cell is adequately discussed in

most chemistry books. The reactions taking place at the carbon electrode are complicated. See reference for details about the reactions.¹

Common type dry cells. Show sections of dry cells, cut vertically.

- b. Mallory mercury-alkaline cell.
- 3. Storage cells.
 - a. The lead storage cell. The reversibility of this cell is the main advantage. The cell has a low internal resistance that makes it useful in starting automobiles.

Storage battery. Construct a storage battery by using strips of lead, strips of plastic, and about 5 M sulfuric acid. The strips of lead and plastic should be about 5 cm wide and 12 cm long. Use plastic strips as separators for the lead plates. Assemble lead plates in beakers. Fill beakers with 5 M sulfuric acid solution. Connect to battery charger for at least ten minutes. Connect storage battery to a bell.

- b. The Edison nickel-iron cell.
- c. The nickel-cadmium cell.

Pupil project—experimenting with nickel-cadmium storage battery. This project, suitable for pupils, and described in *The Science Experimenter*, No. 563, is published by Science and Mechanics Publishing Company, 450 East Ohio Street, Chicago 11, Illinois. *The Science Experimenter* can be purchased locally at magazine stands. Details are given in the article for the construction of a lead-acid and nickel-cadmium battery. Experiments are suggested that enables the pupil to determine advantages and disadvantages of both kinds of batteries.

- 4. Fuel cells.²
 - a. See laboratory for details of a simple fuel cell.

Simple fuel cell. Use solutions of saturat-

ed ferric chloride, 30 per cent KOH and formaldehyde. Place ferric chloride in a porous cup with a carbon electrode. The other electrode, carbon painted with silver paint, is placed in the 30 per cent potassium hydroxide solution. To activate the cell, add several drops of formaldehyde next to silver electrode. Connect the fuel cell to 1½ volt bulb.

Allis-Chalmers Corporation has produced for school use, a simple fuel cell. Charts, diagrams, and theory of cell operations are also supplied.

F. Electrode potentials.

- 1. Measurement of electrode potentials.
 - a. The potential of an electrode depends on the potential difference between the electrode and the electrolyte.

E.M.F. of cells. Experiments involving the E.M.F. of cells should be given to the pupil. Half-cells containing solutions of 0.5 M, copper sulfate, zinc sulfate, and lead nitrate with electrodes of the respective metals should be used.

Each pair of half-cells should be connected with a salt bridge. The pupil should diagram each cell showing direction of electron flow and ion migration in the cell. Equations for each half-reaction should be written. The pupil should calculate the cell voltage from the standard oxidation potential. Then the pupil should measure the cell voltage using a high resistance voltmeter or potentiometer. A salt bridge can be made by bending 5 mm glass tubing into a U-shape. The U-tube is then filled with a solution of potassium chloride. The ends of the U-tube should be plugged with cotton to prevent siphoning of the KCl in case the electrolyte in one half-cell is a little higher than that in the other.

- b. It is impossible to measure directly the potential of a single electrode.
- c. It is more useful to know how the magnitude of an electrode's potential compares with other electrode potentials. The instructor should refer to a college text for

¹ Karol J. Mysels, "The Reaction of the LeClanché Dry Cell," *Journal of Chemical Education* (December, 1955) 638-639.

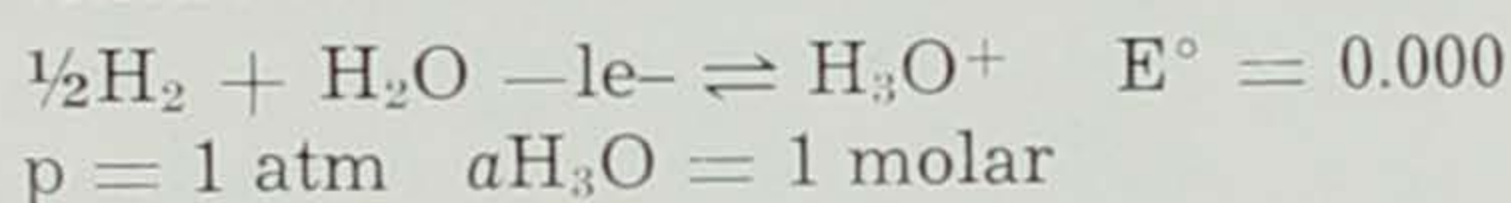
² Leonard G. Austin, "Fuel Cells," *Scientific American*, v. 201 (October, 1959), pp. 72-78.

details on the electrode potential and the oxidation-potential table.

- d. Electrode potentials are compared to a single standard, the hydrogen electrode. See demonstration F.2.a. in this section.

2. Hydrogen electrode.

- a. The reference electrode is assigned a potential of zero.



Electrode potential. Hydrogen electrode. This demonstration gives the necessary details for the determination of the electrode potential of zinc by using hydrogen electrode as the standard.

Hydrogen and chlorine. Electrolysis of a solution of HCl using carbon electrodes for about twenty minutes is first shown. The hydrogen and chlorine absorbed by carbon is sufficient to operate the cell as a primary cell for a short time. The theoretical voltage of this cell is almost produced.

A hydrogen electrode of the Wendt type is easily made. Take a side arm test tube, heat the extreme closed end in a small hot flame and blow out to produce a hole about $\frac{1}{4}$ inch in diameter. If no side arm test tube is available, seal a side arm to a piece of 12 mm tube and then shrink the far end to a $\frac{1}{4}$ inch opening. This makes the outer shell of the electrode as shown in FIGURE 13.

Next, seal a 2 cm piece of platinum wire in 5 mm glass tube so that most of the wire sticks from the end. Connect the wire on the inside of the seal to a long insulated copper wire by means of a globule of mercury or melted Wood's metal. This is the electrode. Plate black platinum on this electrode using it as the cathode in platinizing solution as given in Lange's or other chemistry handbooks. Slip a rubber stopper over this electrode and attach it to the outer shell as shown in the diagram so that half of the platinized platinum wire extends through the opening in the end of the shell. Use part of another stopper as guide.

Rinse the electrode well with water before using and always store under distilled water. Do not allow black platinum

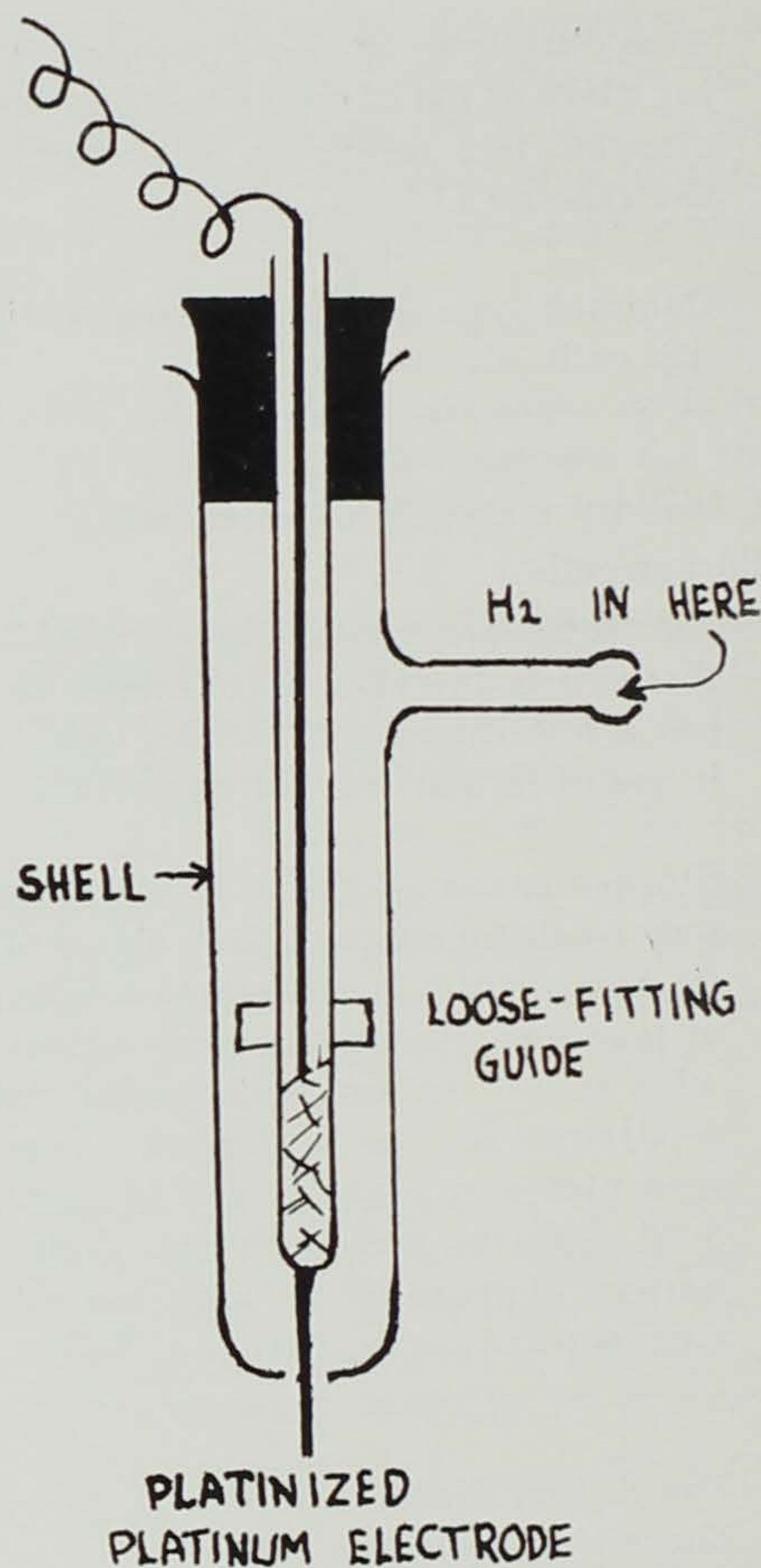


Figure 13

surface to dry or it will have to be replatinized.

To use, immerse the electrode in the solution of HCl or other acid-base solution to be tested for acidity, far enough so that the end of the shell is beneath the surface. Bubble hydrogen gas through the electrode and allow it to escape around the platinum wire at the end. For a standard hydrogen electrode use 1.18 m HCl.

NOTE: Electrolysis of dilute NaOH at nickel electrodes is a good source of hydrogen for hydrogen electrode.

3. Electromotive series.

- a. Development of the electromotive series.
 - (1) The table listing the reactivities of metals and nonmetals relative to each other is known as the electromotive series.

Protection of metals as Fe by other metals as Zn from corrosion.¹ Use of Feroxyl reagent for various demonstrations. W. B. Meldrum, *Journal of Chemical Education*, 25 (1948), 254.

- (2) The sample table as listed in high school chemistry books is inadequate.
- (3) The standard oxidation-potential table can be used profitably by beginning pupils.

4. Standard oxidation-potential table.

Development of an oxidation-potential table. The pupil can develop a simple oxidation-reduction table. Starting with strips of copper, lead, zinc, and solutions containing ions of these metals, a simple oxidation-reduction table can be experimentally developed. The pupil immerses the electrodes in the various solutions and from his observations writes ionic equations for any observed reaction. The results can be tabulated in a table. The table should be arranged with the strongest reducing agent at the top and the weakest reducing agent at the bottom. A suggested arrangement of the table is given below.

Reducing agent	Oxidizing agent	Reaction	E°
Zn	Zn ⁺⁺	$\text{Zn}^\circ = \text{Zn}^{++} + 2e^-$	+0.76V
Pb	Pb ⁺⁺	$\text{Pb}^\circ = \text{Pb}^{++} + 2e^-$	+0.12
Cu	Cu ⁺⁺	$\text{Cu}^\circ = \text{Cu}^{++} + 2e^-$	-0.344

The table can be expanded to include H₂-H⁺, Ag-Ag⁺, Cl⁻-Cl₂, Br⁻-Br₂, and I⁻-I₂.

a. Information provided.

- (1) The table gives the numerical value for each oxidation potential at 25° C when the concentration of the dissolved species is 1 molal.
- (2) The table lists reducing* agents in order of their decreasing strength and lists oxidizing agents in order of their increasing strength.
- (3) A positive oxidation potential indi-

cates that the reduction agent is stronger than hydrogen, while a negative oxidation potential indicates that the reducing agent is weaker than hydrogen.

- (4) The reversibility of the electrode reactions is indicated by the table.
- b. Uses of the standard oxidation-potential table.

The uses of the standard oxidation-potential table can be readily demonstrated by the instructor. For example: The instructor can have the pupils predict the products at the electrodes during the electrolysis of copper chloride solution. Use of the oxidation-potential table in predicting products of electrolysis will be necessary in the next session.

- (1) The table can be used to predict:
 - (a) Products of electrolytic cells.
 - (b) If an oxidation-reduction reaction will take place spontaneously.
 - (c) The direction of electron flow in a voltaic cell.
 - (d) The voltage of a cell.
- (2) The instructor should emphasize that a reducing agent will reduce any oxidizing agent below it in the table. An oxidizing agent will generally oxidize any reducing agent above it in the table.

NOTE: These rules are true only if the ions are about equal in concentration. The ion concentration in the standard oxidation-potential table is 1 molar. If the concentrations of ions are unequal, there may be a shift in the position of the oxidizing power. In this unit it will be better not to complicate the situation, only assume the ion concentrations of 1 molar.

G. Electrolysis.

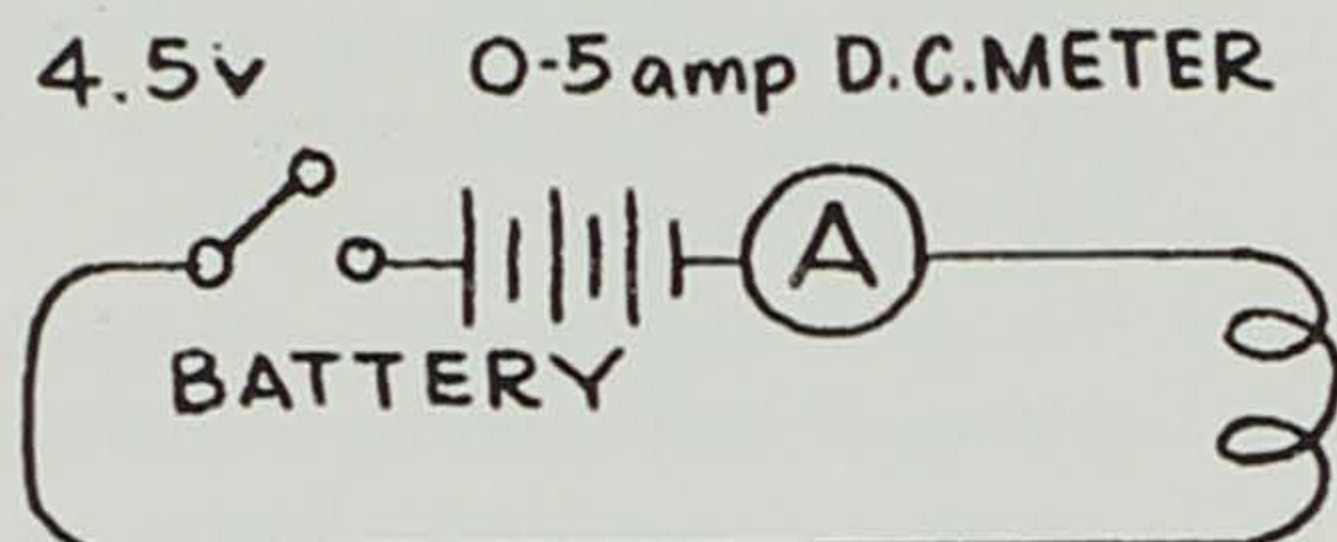
1. Metallic conduction.

Nature of metallic and nonmetallic conduction. Connect a 2-foot length of resistance wire, e.g., 20 gauge nichrome wire, in the circuit shown below (FIGURE 14). Coil the

¹ Manufacturing Chemist's Association, Inc., "Corrosion of Iron—Experiment No. 24," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 47.

* NOTE: Some tables use a different convention. Instructor should examine table used by class to determine the convention used and whether strongest reducing agent is at the top.

wire so it can be heated but so that the coils do not touch.



COIL RESISTANCE 2-3 ohms

Figure 14

Close the switch and observe the current which flows. Does the coil warm up? How does the current change as the coil warms? Heat the coil to redness with a Bunsen burner. How has the rise in temperature affected the resistance of the coil and hence the current passing through it?

Repeat the experiment with a $\frac{1}{4}$ inch carbon rod 6 to 8 inches long in place of the nichrome wire. How does a rise in temperature affect the resistance of the carbon?

Repeat, using an electrolytic cell.

-
- a. The instructor should compare metallic conduction with electrolytic conduction.
 2. Electrolytic conduction.
-

Electrolytic conduction of hot glass. Wind heavy copper wire of 12-14 gauge tightly around the ends of a piece of $\frac{1}{4}$ inch soft glass rod about 5 inches long. Connect to 115 V circuit through a ballast resistor capable of passing 5 amperes or more such as an old electric flatiron or hot plate. Heat the glass to red heat with a burner and the temperature will remain high and become higher due to electrolytic conduction by the glass. The heat can be controlled by moving the ends of the glass closer together or farther apart.

- a. Use the diagram below to identify and discuss the meaning of electrolyte, electrode, anode, cathode, anions, and cations.
 - b. Use these equations to discuss reactions at the electrode.
 - (1) Anode reaction.

$$A^- \rightarrow A^0 + e^- \quad (\text{oxidation})$$
 - (2) Cathode reaction.

$$C^+ + e^- \rightarrow C^0 \quad (\text{reduction})$$
-

NOTE: The cathode is always the electrode at which electrons enter the cell from an outside conductor, hence it is the electrode at which reduction takes place. This is true both for primary cells and electrolysis cells.

3. Qualitative nature of electrolysis.

- a. What is formed at the electrodes? Electrons leave the anode and travel through the external circuit to the cathode. A chemical change must take place at the anode that will supply these electrons. Oxidation is a reaction that involves the loss of electrons. Therefore, oxidation takes place at the anode. One can follow a similar reasoning to show that reduction occurs at the cathode.

b. Electrolysis of fused NaCl.

The discussion of what happens when a current is passed through fused NaCl is recommended. Discuss the electrode reactions in terms of oxidation and reduction.

c. Electrolysis of aqueous solutions of salts.

- (1) During electrolysis, three reactions are possible at the electrodes. The electrodes may react, the water may gain or lose electrons, or the ions of the solute may gain or lose electrons.
-

Electrolysis of aqueous solution of zinc chloride, zinc sulfate, and potassium iodide should be carried out by instructor. A U-tube fitted with carbon electrodes is satisfactory. Indicators can be added to the U-tube to help identify products of electrolysis. The oxidation potential table should be used to predict the products of electrolysis before the demonstrations are performed.

The pupil should be given the opportunity of doing some electrolysis experiments. Three U-tubes fitted with carbon electrodes should be provided for each pair of pupils. Solutions of sodium bromide, sodium sulfate, and copper chloride can be used. The pupil should predict the electrode products before doing the actual electrolysis.

- (2) It is possible to predict which one of the three possible reactions will take place if one knows which reactant is the best oxidizer or reducer.
 - (3) The best oxidizer or reducer may be predicted from the oxidation potential table developed earlier in this unit.
 - (4) The reaction at the cathode is not easy to predict. The best oxidizer is not always reduced since factors such as the rate of reduction, concentration of the reactant, and current effects must be considered. The following rules may be applied in many reactions. If the solution contains ions of inactive metals, generally the metal ions will be discharged. If the solution contains active metal ions, generally water molecules or hydrogen ions will be discharged.
 - (5) The best reducer is generally oxidized at the anode. The following rules may be used to predict reactions at the anode. If the anode is made of a metal that is a better reducer than the anion, the anode will go into solution as positive ions. If the electrode is inert and the anion is a better reducer than water, the anion will generally be discharged. If the electrode is inert and the anion contains non-metal atoms in their highest oxidation state, the water molecule or the hydroxyl ion will be discharged.
 - (6) Use the demonstration to show applications of the general rules. Formation of metals from their ions, oxidation and reduction of water, and discharge of a simple anion are shown.
4. Quantitative nature of electrolysis.
- a. How much is formed at the electrodes?
 - b. Electrical units.
Discuss the meaning of volt, ampere, ohm, coulomb, and faraday.
 - c. Faraday's laws of electrolysis.
 - (1) The quantity of substance produced at the cathode and anode during electrolysis is proportional to the quantity of electricity passed through the cell.

— — — — —
Either of the experiments listed be-

low can be an instructor demonstration. The instructor can do the experiment during a discussion of the Faraday's Laws of electrolysis. Pupils can use the data acquired to obtain the value of the faraday, equivalent weight of copper, electrochemical equivalent of copper, or mass of a copper atom.

Determination of the faraday. Electrolysis of a solution of sulfuric acid using a platinum electrode as the cathode and a copper anode is the basis of this experiment.

The hydrogen evolved is collected in a burette tube. The copper anode is weighed before and after electrolysis to obtain the amount of copper oxidized. By determining the amount of copper dissolved and the volume hydrogen produced during a measured time with a known current, the value of the faraday can be determined. The experiment, "The Number of Atoms in a Hydrogen Molecule,"¹ can be used to obtain the value of the faraday. This experiment is suitable for schools which do not possess an analytical balance. The simplicity of this experiment makes it suitable for either a pupil experiment or an instructor demonstration.

- — — — —
- (a) The instructor should develop a mathematical expression of the first law. Problems should be given to the pupil to help him understand the law.
 - (b) The laboratory experiment is designed to follow the discussion of the first law. The value of the faraday and the equivalent weight of copper can be determined.
- (2) The amount of different substances produced at the cathode or anode by the same quantity of electricity is directly proportional to their equivalent weights.

¹ Manufacturing Chemist's Association, Inc., "The Number of Atoms in a Hydrogen Molecule—Experiment No. 3," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 5.

— — — — —
 A diagram that shows relative amounts of different substances produced during electrolysis helps explain this law. A series of electrolytic cells hooked in series should be used.
 — — — — —

- (a) The instructor should develop a mathematical expression of the second law.
- (b) Simple problems using this law should be given to the pupil.

UNIT XII

ORGANIC CHEMISTRY OF CARBON

Carbon atoms can bond in an unusual way. This gives rise to many different compounds having the same kind of atoms.

A. Bonding and structure.

— — — — —
 Show models of the carbon atom. Review the concept of covalence. Discuss role of hydrogen as a hydride. Put concentrated sulfuric acid into a beaker of sugar. Use just enough acid to completely dampen the sugar. Wait. Have the beaker sitting on a large plate to protect the table top. Call attention to carbon as the organic skeleton.

Explain ground and excited states of atoms to pupils. Review mental picture of orbitals.

Demonstrate the difference between organic solvent, as CCl_4 , and water. Compare the physical properties of the two substances.
 — — — — —

1. Carbon combines with four other atoms to form three dimensional structures of many kinds. Example:

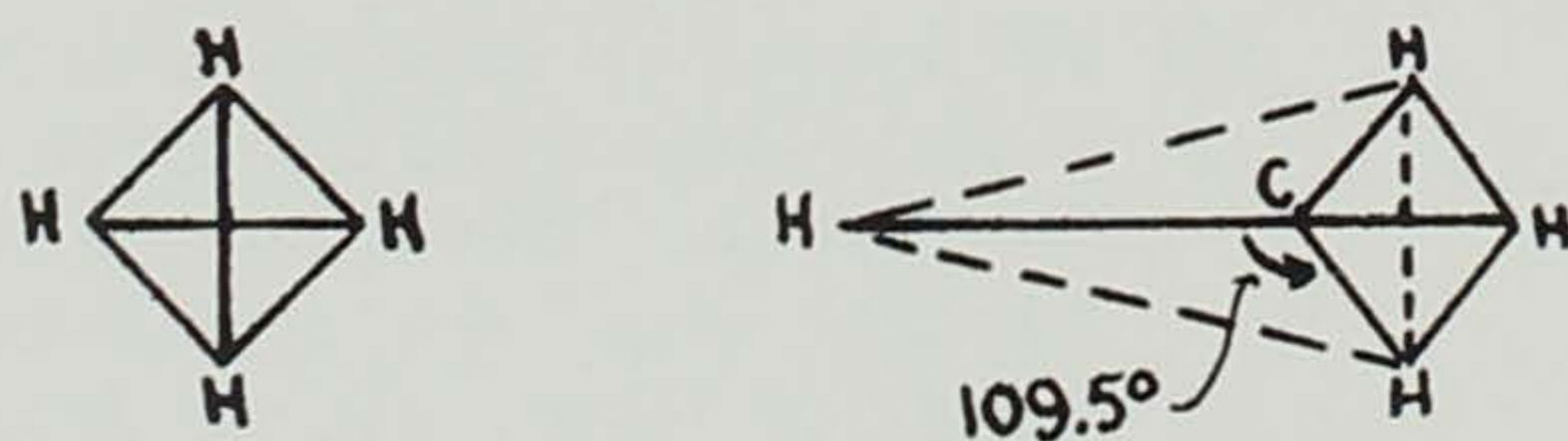


Figure 15

Film: *Carbon and Its Compounds*. S.U.I. No. U-2683, 11 min., b & w. Sound. Also, I.S.U. NS-3390, 11 min., b & w.

Film: *The Waiting Harvest*. 525 William Penn Place, Pittsburgh 30, Pa.: U.S. Steel Co.

Film: *Modern Miracle Makers*. Manufacturing Chemist's Association, Inc. (Note: Send for

their complete film guide. Address: 1825 Connecticut Ave., Washington 9, D.C.)

2. The carbon atom has two 1s electrons, two 2s electrons, and two 2p electrons. Example:

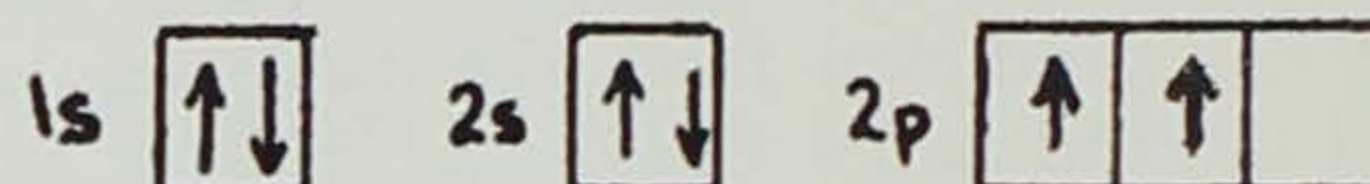


Figure 16

3. When combining with other elements, one of the 2s electrons moves from the ground state to the excited state and occupies one-half of a 2p suborbital. This gives carbon its characteristic oxidation number of four, with one 2s and three 2p orbitals open for sharing electrons from other atoms. Example:

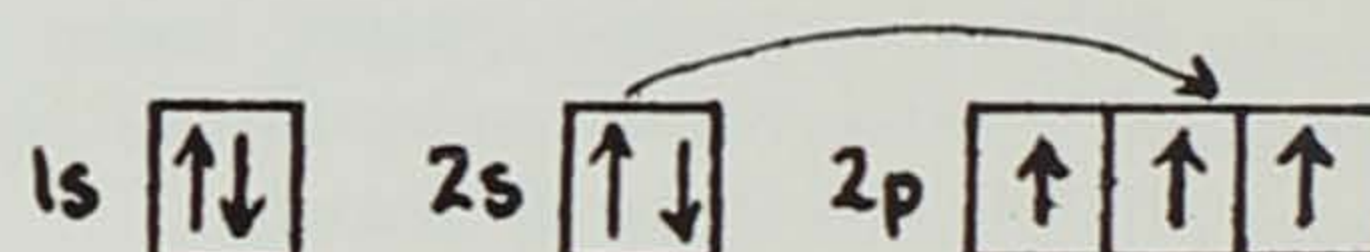
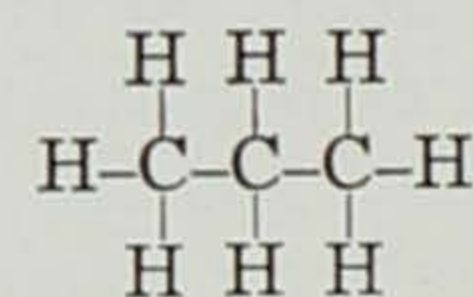
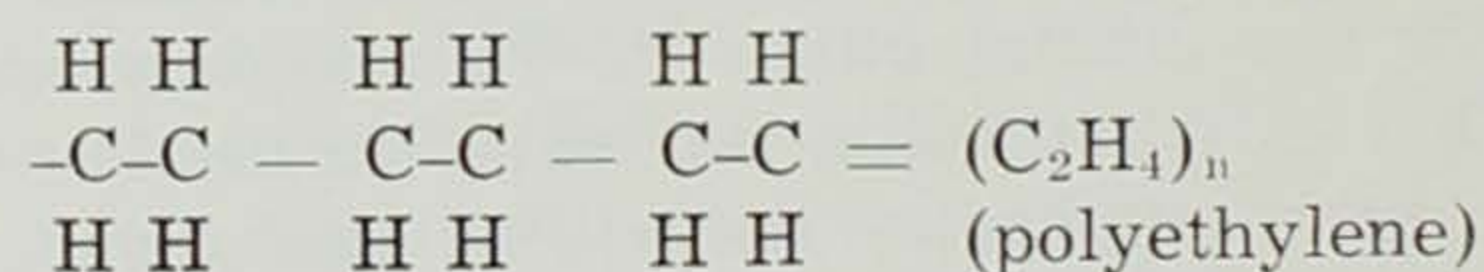


Figure 17

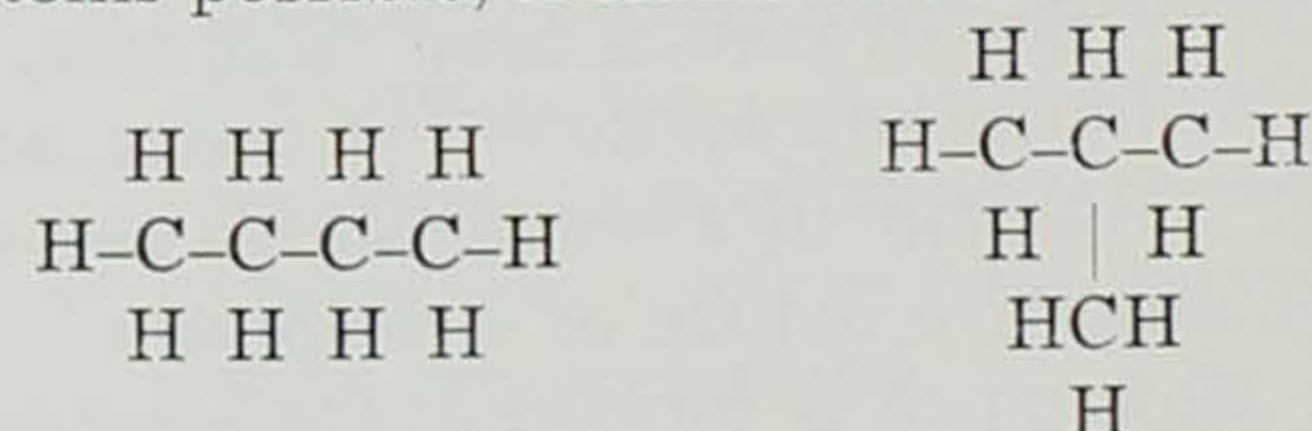
4. Carbon combines with hydrogen relatively easy forming the typical hydrocarbon structure.



5. Because it can bond to itself, it forms polymers which have characteristic properties.



6. Because of different arrangements of carbon atoms possible, it forms structural isomers.



B. Classes of organic compounds—nomenclature.

— — — — —
 Draw several hydrocarbon structures. Suggest reasons why oxygen might become involved in hydrocarbon chains. Note reducing properties of both hydrogen and carbon. Introduce concept of polymers.

Give practice in naming hydrocarbon chains. Show how the bonds can be compounded to form unsaturated chains.

Mention that there are many other types of saturated and unsaturated hydrocarbons that have not or cannot be discussed in the time allotted. Suggest at this point that the field of organic chemistry is in its infancy and is tremendously complex even at this early date.

Hydrocarbons may be classified as saturated or unsaturated. Unsaturated aliphatic hydrocarbons *add* Br₂ readily from CCl₄ solution. Saturated hydrocarbons do not react with Br₂ in the cold. Unsaturated aromatic (ring) compounds react with Br₂ by substitution of Br for H with the evolution of HBr which fumes when the breath is blown over the container.

Test each of the following substances with a 5 per cent* solution of Br₂ in CCl₄ by adding the bromine solution drop by drop and shaking until the brown color of free Br₂ persists: pentane, benzene, phenol, kerosene, gasoline. Test natural gas and acetylene by bubbling the gases through 3 ml portions of the Br₂ solution in CCl₄. Make acetylene by adding water to calcium carbide in small flask or test tube fitted with stopper bearing a medicine dropper and a delivery tube. Classify the substances tested according to the results of this test.

An alternate test for unsaturation is that of Baeyer using potassium permanganate solution. Unsaturated aliphatic compounds will decolorize a considerable volume of 0.1 N KMnO₄ solution while saturated compounds, not reducing agents, will not. Reducing agents such as aldehydes give the Baeyer test even though saturated. The instructor should consult a good organic chemistry text in doubtful cases. Also note that traces of reducing impurities will decolorize the first few drops of permanganate solution.

1. Hydrocarbons—compounds containing only carbon and hydrogen.

- Alkane series or paraffins with type formula C_nH_{2n+2} —methane, ethane, propane, butane, pentane, hexane, etc.
- Alkene series or olefins with type formula C_nH_{2n} —ethene, propene, butene, pentene, hexene, etc.
- Alkyne series or acetylenes with type formula C_nH_{2n-2} —acetylene or ethine, propyne, butyne, pentene, hexene, etc.

* Use extreme care in handling liquid bromine. Do not let pupils make up the 5 per cent solution. The solution, once made, is safe for pupil use.

d. Aromatic series or benzene series with ring structures and formulas C₆H₆—benzene, C₁₀H₈—naphthalene, C₁₄H₁₀—anthracene.

e. Other less common series.

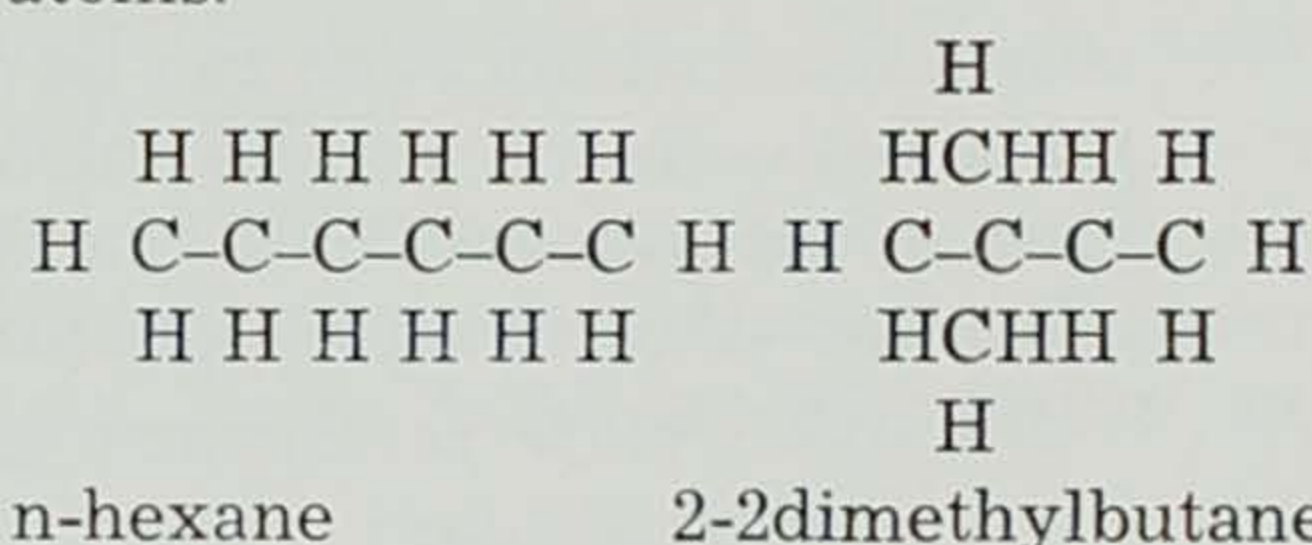
f. Hydrocarbons may arrange themselves in various ways resulting in structural differences but with identical empirical formulas. See isomers, XII.A.6, page 54.

— — — — —
Show the various types of structural formulas. Give practice in writing and interpreting these formulas.
— — — — —

Film: *Refining Oil for Energy*. Shell Oil Co., 624 So. Michigan Avenue, Chicago 5, Illinois.

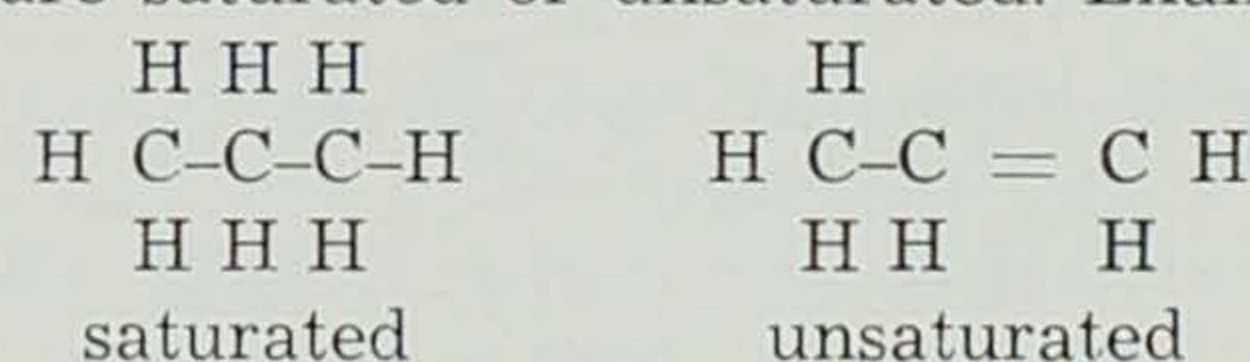
Film: *Oil for Aladdin's Lamp*. (Chemicals from Petroleum.) Shell Oil Co.

- (1) We use structural formulas to denote the difference in the arrangement of atoms.



- (2) We use different names to denote structural differences. Examples above.

g. Hydrocarbons are further separated by their bonding according to whether they are saturated or unsaturated. Examples:



— — — — —
Discuss the difference between saturated and unsaturated hydrocarbons. Suggest why the difference in bonding results in different properties and uses.
— — — — —

2. Derivatives of hydrocarbons. Radicals can replace hydrogen on already saturated hydrocarbons or can be added to unsaturated hydrocarbons forming numberless compounds called derivatives.

— — — — —
Methods of preparation for the common organic compounds should be discussed in class

and the mechanisms involved in the reactions clarified where they are simple enough to be easily understood by pupils with little chemical background. Complex organic mechanisms had better be left for advanced courses in organic chemistry.

It is interesting and often helpful to pass small samples of common organic chemicals around a class while calling attention to their often peculiar odors and other significant properties.

Where possible, the common uses of the various compounds discussed should be emphasized, e.g., benzaldehyde as artificial almond flavor and methyl salicylate as wintergreen flavor.

Pupils should be encouraged to read the current issues of *Chemical and Engineering News* for information on recent methods for production of organic compounds as well as new uses for them. A number of simple syntheses can be made as lecture demonstrations. This is especially true of esters. In most cases esterification can be made to take place between the acid and alcohol in a test tube in the presence of a few drops of concentrated sulfuric acid. Examples are:

Amyl acetate (acetic acid and
amyl alcohol)
Methyl salicylate (salicylic acid
and methyl alcohol)

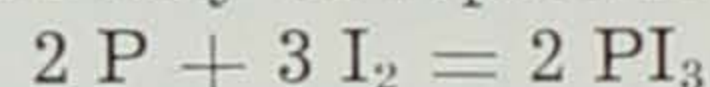
Others may be tried as ingredients are available. Many of the laboratory manuals give more specific directions.

a. Halogenated derivatives, alkyl halides
 RX_n . Example:



Prepare a typical alkyl halide, ethyl iodide. Materials: 2.5 grams red phosphorus, 11.6 cc (9.2 g) 100 per cent dry ethyl alcohol, 25.4 g iodine. Unless 100 per cent or absolute alcohol is available, first prepare this from 95 per cent material by drying over quick lime, CaO, for two days. Distill the dry alcohol from the mixture in a system protected from moisture by drying tubes containing $CaCl_2$. Use a hot water bath as a source of heat. Directions: Attach a reflux condenser to

a 200 ml flask by means of a tight fitting cork and support the assembly firmly 6 inches above the desk top. Add directly to the flask 2.5 g of red phosphorus and 11.6 cc of the absolute alcohol. Then, in portions of 2 grams each, add 25.4 grams of coarsely powdered iodine, reattaching the condenser and shaking the flask after each addition of iodine. If the mixture gets very hot and is in danger of boiling, cool it by holding a pan of cold water around the flask. The reaction which occurs may be represented as:



When all of the iodine has been added, allow the reaction mixture to stand for twenty-four hours or until the next laboratory period. After standing, heat on a water bath for thirty minutes, cool, replace the reflux condenser with one arranged for distillation, and distill the volatile ethyl iodide into a small flask cooled with ice or cold water and loosely stoppered. Use a very low flame and distill slowly. Place the distillate in a separatory funnel if one is available or in a graduate. Add some water, shake, and test with litmus. If acid, neutralize with dilute sodium hydroxide. If yellow, add dilute sodium thiosulfate solution until colorless. Allow the treated mixture to stand and it should separate into two layers. Care! The ethyl iodide is quite volatile and should be prevented from evaporating. Pour off the upper or water layer and separate as carefully as possible from the lower or ethyl iodide layer. If the product still does not appear pure, treat again with NaOH, $Na_2S_2O_3$ solutions and water and again separate the heavy bottom layer of ethyl iodide. Remove the last traces of water with anhydrous calcium chloride solid letting the product stand in contact with the calcium chloride for several hours. Filter through glass wool and distill. The ethyl iodide should boil at 72.2° at 760 mm pressure.

In addition to the above typical organic preparation, there are many more simple enough for high school classes which require only common reagents. Among these are: synthesis of urea from ammonium cyanate, preparation of aspirin, and

the synthesis of methyl orange. Directions for these are to be found in teacher's edition with instruction manual.¹

Preparation of alcohols, aldehydes, and acids.

Preparation of esters.

Properties of sugars and carbohydrates in general.

Properties of proteins.

Preparation of a soap and study of detergents.

Properties of textile fabrics.

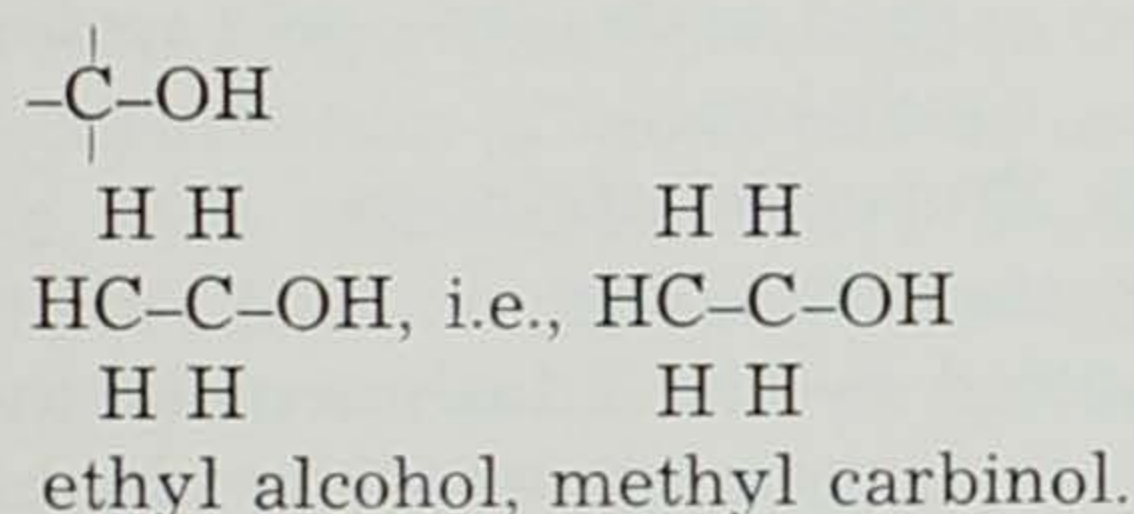
Directions for these are in Young and Petty, *Chemistry for Progress Laboratory Manual*. Englewood Cliffs, New Jersey: Prentice-Hall, Inc., 1957.

Experiment on rubber.²

The following experiment using the ethyl iodide made above, UNIT XII.B.2.a., p. 56, is given in the reference, "The Grignard Reagent Reaches the Freshman."³

Ethyl iodide works as well as the methyl iodide originally called for.

b. Alcohols (carbinols)—carbinol group is



At this point call attention to the fuel properties of the hydrocarbons. Burn alcohol and suggest what the products might be. Burn acetylene and try to get the pupils to reason out the difference in combustion products. Point out that soot is free carbon, and let them take it from there. Recall attempts to use alcohol-gasoline mixtures as automobile fuel.

Discuss the necessity of knowing the chemical nature of alcohols before use. What effect does the length of the hydrocarbon chain have on properties and uses? What effect does the position of the

OH group have on properties and uses? If a variety of alcohols are available, have pupils smell of each and compare with structure.

Stress the poisonous nature of methyl alcohol and precautions which should accompany its use, especially as a radiator antifreeze. Recall the almost complete shift from methyl alcohol as a product of wood distillation to modern synthetic production.

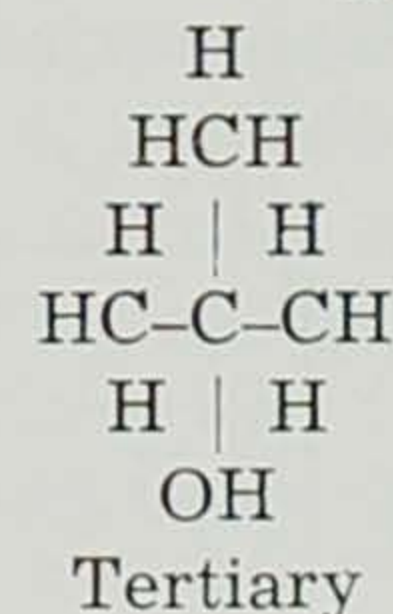
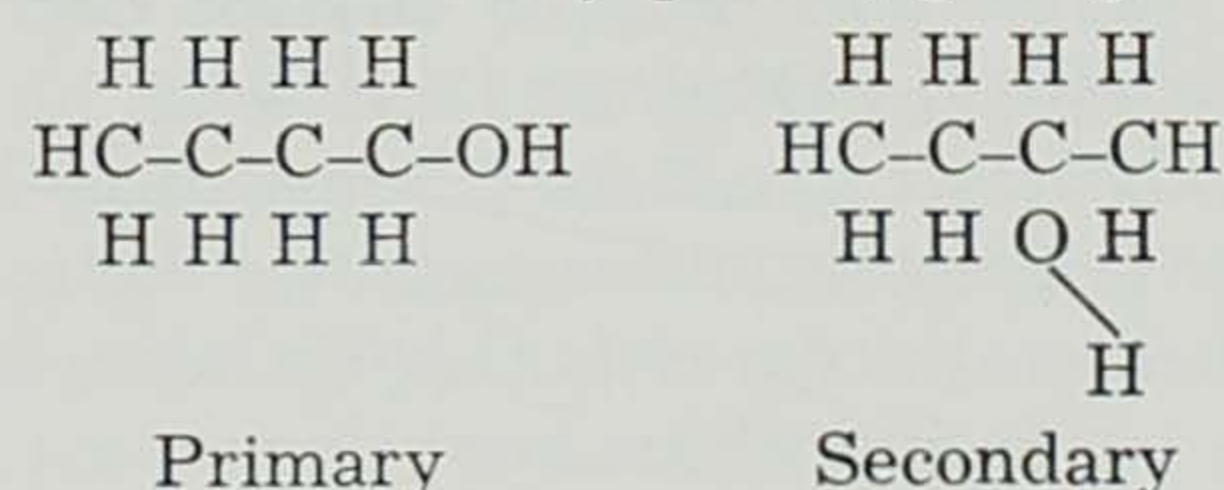
Discuss the tax structure on ethyl alcohol and the reasons and methods for "denaturing" ethyl (grain) alcohol. See chemistry handbooks for denaturing agents and approved formulas for this.

Mention isopropyl alcohol as cosmetic and rubbing alcohol.

Film: *The Man in the Doorway* (organic chemistry). American Cyanamid Co. Distributed by Modern Talking Picture Service, 216 E. Superior St., Chicago 11, Illinois.

Film: *Welcome to Hercules* (activities and products in chemical research). Hercules Powder Co., Delaware Trust Building, Wilmington 99, Delaware.

- (1) Alcohols have a type formula R-OH where R represents the alkyl or hydrocarbon chain.
- (2) Alcohols are named from the corresponding hydrocarbons. As methyl-, ethyl-, propyl-, butyl-, amyl-, and hexyl- after which they take the Greek numerical prefixes.
- (3) The OH radical may be added or substituted in many places, giving:



- (4) Preparation of alcohols (many ways).
 - (a) Methyl alcohol by direct syn-

¹ Harper W. Franz and Lloyd E. Malm, *Essentials of Chemistry in the Laboratory* (San Francisco: W. H. Freeman and Company, 1961).

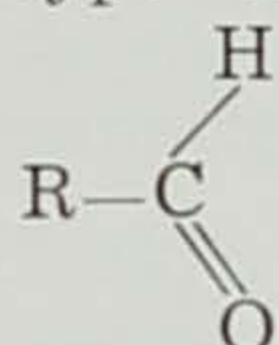
² Manufacturing Chemist's Association, Inc., "Experiment on Rubber—Experiment No. 29," *Scientific Experiments in Chemistry* (New York: Holt, Rinehart and Winston, 1959), 57.

³ W. B. King and J. A. Beel, "The Grignard Reagent Reaches the Freshman," *Journal of Chemical Education*, XXXII (1955), 146.

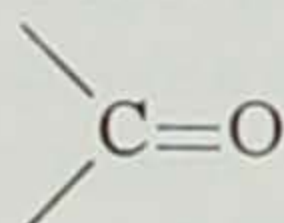
thesis from CO and H using a catalyst.

- (b) Ethyl alcohol by fermentation.
- (c) By hydration of unsaturated hydrocarbons.
- (d) By hydrolysis of alkyl halides.
- (e) Other methods at the discretion of the instructor.

c. Aldehydes have a type formula.

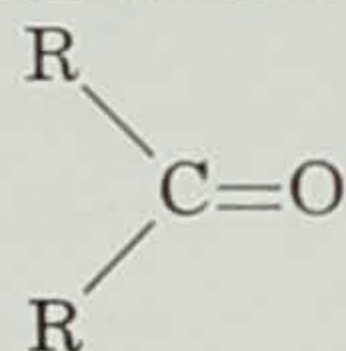


- (1) Contain the carbonyl radical.



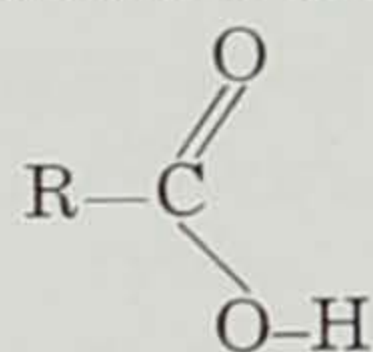
- (2) Preparation by oxidation of a primary alcohol or by many other methods.
- (3) Properties and uses.

d. Ketones have a type formula.



- (1) Also contain the carbonyl radical but with *two* alkyl groups attached.
- (2) Preparation by dehydrogenating a secondary alcohol or by other methods.
- (3) Properties and uses.

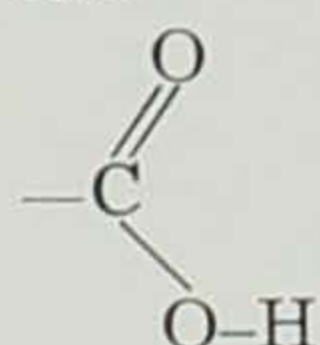
e. Acids have the type formula.



Films: Four films varying from 11 to 20 minutes on *Techniques of Organic Chemistry* by Professor Louis F. Fieser are available from S.U.I. on a rental basis. Part I, U-4496; Part II, U-4497; Part III, U-4498; Part IV, U-4499.

These are excellent, but are on *college* or *adult* level. They should be seen by the instructor before being ordered for high school class use. May be too technical for most classes.

- (1) Characterized by the carboxyl radical.

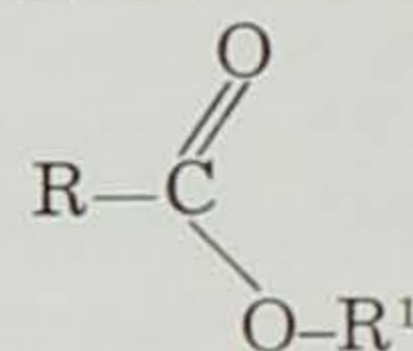


- (2) Preparation from alcohols or aldehydes by oxidation and by other methods.

- (3) Carboxylic acids are weak acids, i.e., are incompletely ionized in water solution.

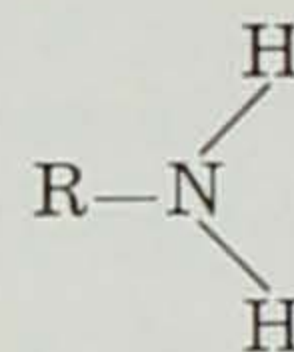
- (4) Properties and uses.

f. Esters have a type formula.



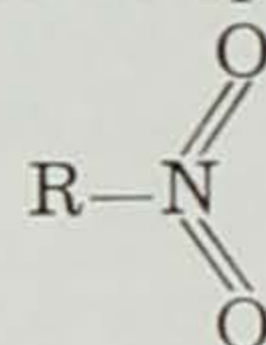
- (1) Preparation from an organic acid and an alcohol with removal of water.
- (2) Properties and uses.

g. Amines have a type formula.



- (1) Preparation by reaction of ammonia with an alkyl halide.
- (2) Properties and uses.

h. Nitro compounds have a type formula.



- (1) Preparation by nitration of a hydrocarbon with nitric acid under proper conditions.
- (2) Properties and uses.

3. Aromatic derivatives.

- a. Formed by the attachment of the above characteristic groups on hydrocarbon rings or skeletons having both rings and chains of carbons.
- b. Basic compounds obtained from coal tar and from certain petroleum sources.
- c. Many of our most valuable dyes, medicines, plastics, and other useful products from this group.

— — — — —

Manufacturing Chemist's Association, Inc., "Preparation of DDT—Experiment No. 31," p. 61. Note: An *excellent experiment* illustrating synthesis of a useful product. Requires close supervision of pupils for good results and safety. Necessary precautions are given in the directions.

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- d. Specific compounds or groups may be considered at the discretion of the instructor.

C. Some applications of organic chemistry.

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Heat some protein substances in a test tube having pieces of litmus paper and lead acetate

paper over the lip. Blackening of the protein shows its carbon content, litmus turns blue indicating ammonia, lead acetate paper turns black if S is present indicating the presence of hydrogen sulfide and water condensed on the side of the tube.

The *Journal of Chemical Education*, Volume 36, 1959, has a number of pages of Tested Demonstrations which can also be adapted to suitable laboratory experiments if the factor of safety is considered in making the proper choices. Page numbers are as follows: (1) industrial organic products, page A234; (2) dyes, plastics, perfumes, pages A297-298; and (3) soaps, organic preparations, pages A377-378.

Film: *Rubber in Today's World*. S.U.I. U-4749.

Film: *Cotton, From Seed to Cloth*. S.U.I. I-1163, 1164.

Film: *Synthetic Fibers*. (Nylon and Rayon). S.U.I. U-2926. 16 min.

Film: Series of Three. *Origin of Life*. Aeronautics and Space Administration, Washington, D.C.

1. Fuels.

An interesting demonstration on the destructive distillation of wood and of coal is given in a high school laboratory course by DeBruyne, Kirk, and Beers; *Semimicro Chemistry*; Experiment 21. They point out that more than 200,000 compounds are made from derivatives of soft coal obtained during the cooking process. The coal is distilled in a hard glass test tube with a drying tube connected by means of two rubber stoppers and a short piece of glass tubing as a condenser. The liquid products collect in the bulb of the drying tube, and gaseous products escape from the small tube at the far end. These are tested for acids with litmus and for hydrogen sulfide with lead acetate paper.

2. Textiles, plastics, and elastomers.
3. Foods and nutrition.

The chemical nature of all foods should be stressed. How far the complex processes of digestion of carbohydrates, fats, and proteins should be pursued in a general course of this kind must be left to the individual instructor.

The importance of the many food additions now used in the preparation and preserva-

tion of commercially marketed food products should be mentioned.

Essentials of Chemistry in the Laboratory by Franz and Malm mentioned earlier gives simple tests for the common food constituents that can be used in a demonstration.

- a. Proteins.
 - b. Carbohydrates.
 - c. Fats.
 - d. Body chemistry.
4. Drugs, dyes, etc.

Dyeing of textiles makes an interesting and colorful lecture demonstration. Commercial dyes can usually be obtained locally. It is wise to try these on the samples of cloth to be used in the demonstration before going before a class as results are not always sure with modern treated cottons, wools, etc.

The distinction between colored chemical substances and dyes should be made, i.e., dyes attach themselves to or react with the fiber of the cloth so that the color is more or less "fast" to washing, etc. The use of mordants in dyeing may be mentioned. This furnishes a good opportunity to show the interrelation of inorganic chemistry to organic chemistry.

"Rapid Preparation of 6-6 Nylon."¹

Place the heavier solution in the beaker with the lighter solution on top. Pick up the film which forms between the two layers, and pull so as to remove a continuous thread.

Preparation of an organic dye, Fluorescein. Mix 0.1 g each of phthalic anhydride and resorcinol in a test tube and add three to four drops of concentrated sulfuric acid. Heat gently for several minutes. Allow to cool, add 5 ml of water and make alkaline with sodium hydroxide. Transfer a drop of this solution to a test tube full of water and view both by reflected and transmitted light. The phenomenon of fluorescence explains the name of this dye. Larger quantities can be made and the product purified.

Preparation of an acid-base indicator, Phenolphthalein. Mix 0.1 g each of phthalic anhydride and phenol (CARE: CORROSIVE) in a test tube and add two drops of concen-

¹ Kinsinger, and others. "Rapid Preparation of 6-6 Nylon," *Journal of Chemical Education*, XXXV (1958), A607. (Chem. Ed. Tested Demonstration.)

trated sulfuric acid. Heat gently over a low flame with steady shaking until the mass melts and becomes dark red. Cool, add water, and neutralize with dilute NaOH.

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UNIT XIII

NUCLEAR CHEMISTRY*

A. Historical

1. Perrin showed that cathode rays consisted of negatively charged particles in 1895.
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Use a cathode ray tube and magnet to show nature of cathode rays. A canal ray tube and a magnet can be used to show the nature of canal rays (positively charged ions).

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2. Becquerel, attempting to find the relation between X-rays and fluorescence, accidentally discovered radioactivity in 1896.
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Expose a wrapped piece of photographic film to a radioactive rock for several days. Develop and show to class. Repeat with a dish of orange Fiesta ware. The orange glaze contains uranium.

Radioactivity Experiments for High Schools Using Glazed Ceramics: This experiment by Helen Crawley appears in the *Journal of Chemical Education*, Vol. 36, pp. 202-203. Glazed orange Fiesta ware was used for the series of experiments. Details for using the glazed ceramic for taking autoradiographs are given.¹

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3. J. Thomson measured the ratio of charge to mass of an electron in 1897.

Film: *Electrons in Gases, Liquids and Vacuum*. Encyclopedia Britannica Films, Inc., 11 min.

4. The Curies discovered radium in 1898.
5. Max Planck advanced the quantum theory in 1900.
6. Crookes showed that radioactivity was not limited to the uranium and its decay products in 1900.
7. Einstein advanced the idea $E = mc^2$ to de-

scribe the relationship between energy and mass in 1905.

Note: J. Klinger, Scientific Apparatus, 82-87 160th Street, Jamaica, New York, are suppliers for "Leybold" producers of physics equipment in Germany. The Leybold Company has developed many experiments using the Braum and Wulf electroscope. A manual of experiments for introduction to Atomic and Nuclear Physics is available. The experiments deal with photoelectric effect, ionization, range of alpha particles, absorption of alpha, beta and gamma radiation, inverse square law, half-life of a thorium emanation, Geiger-Muller counter, and the Wilson cloud chamber. The apparatus available from Leybold enables the instructor to develop the fundamental ideas with a minimum of effort and equipment.

B. Nature of radioactivity

1. Three types of radiation from natural radioactivity were recognized by 1900.²
- — — — —

Demonstrate the activity of several radioactive materials using a Geiger-Muller counter.

— — — — —

2. Wave theory of light³
- — — — —

Demonstrate photoelectric effect.

— — — — —

- a. The nature of wave motion.
 - b. Characteristics of transverse waves.
 - c. Electromagnetic waves.
- — — — —

Use two Braum or similar electroscopes equipped with amalgamated zinc plates on top. Charge the zinc plates with charges of the opposite sign. Expose both plates to the light of a mercury vapor lamp. Note that the ultraviolet discharges only the electroscope with the negative charge.

Repeat demonstration and obtain the time required to discharge the negative electroscope. Double the distance between the mercury vapor lamp and the electroscope. The time required to discharge the electroscope should be four times greater. The

* See *Rules for Handling Radioactive Materials* at the end of Unit XIII.

¹ Frank Starr, "Radioactivity, Demonstrations, Experiments and Techniques Using Radioisotopes in Science Teaching," unpublished Master's thesis (Cedar Falls: State College of Iowa). This thesis gives complete details on the use of radioactivity in high school science courses. Experiments, demonstrations along with techniques and materials are given.

² *Nuclear Science Teaching Aids and Activities* (Washington: Department of Health, Education and Welfare, June 1960), pp. 9-30. (reprint)

³ Charles E. Dull; H. Clark Metcalfe; and John E. Williams, *Modern Chemistry* (New York: Henry Holt and Company, 1958), 557-584.

relationship between the intensity of the light and the distance may be developed. Note that the inverse square law holds strictly only for a point source of light. If the light source is a long tube, the tube should be masked by a sheet of metal with a $\frac{1}{2}$ in. hole for the light to pass through. CARE! Ultraviolet light is dangerous to the eyes. Workers should wear glasses or preferably goggles and not look directly at the light. Read directions with the light source. Insert a glass plate and a sheet of plexiglass in turn between the light source and the negatively charged electroscope. Account for the failure to discharge the electroscope when the glass plate is between the light source and electroscope.

3. May be included at the discretion of the instructor. Quantum theory of radiation.
 - a. Wave-particle duality.
 - b. The uncertainty principle of Heisenberg.
 - c. Wave mechanics.
4. Elements possessing natural radioactivity.¹
 - a. Polonium (Po)
 - b. Radium (Ra)
 - c. Actinium (Ac)
 - d. Thorium (Th)
 - e. Uranium (U)
 - f. Radon (Rn)
 - g. Protactinium (Pa)

Film: *Atomic Radiation*. E.B.F., 1956. Explains the fundamentals of atomic radiation role of beta, alpha, gamma; the neutron in research is also discussed.

5. Characteristics of radioactive elements.
 - a. Kinds of rays emitted.
 - b. Half-life of a radioactive element.

Determine the half-life of I^{131} or use the results obtained from *The Science Teacher*, December 1958, p. 442, to plot a graph. From the graph determine the half-life.

6. Kinds of nuclear reactions as to type of particles ejected.
 - a. Nucleus has too high neutron-proton ratio—will either eject a neutron or a beta particle. Beta emission is the most common.
 - b. Nucleus has too low neutron-proton ratio.

Will either absorb a K-orbital electron or emit a positron.

- c. Nucleus has too many protons. Will eject alpha particle and beta particle through a series of steps to form a stable nucleus.

C. Disintegration reactions

(Example: ${}_{92}\text{U}^{238} \rightarrow {}_{90}\text{Th}^{234} + {}_2\text{He}^4$)

1. Natural radioactivity.²

Discuss the natural decay series. Show a chart that gives the branching, type of radiation chain and half-life of the members of the decay series.

- a. Three natural disintegration chains, known as the uranium, thorium and actinium series, exist.
- b. The final decay product of each series is an isotope of lead.
- c. A neptunium decay series also exists; the end product is an isotope of bismuth.
- d. Using the half-life of the members of the natural radioactivity series and the analysis of the decay products, the age of the earth has been estimated to be around 3 billion years.³

Film: *Radioactive Series*. McGraw-Hill Book Company Film Series, b&w, 1954.

Film: *Transuranic Elements*. S.U.I.

2. Induced radioactivity.⁴

- a. Artificial radioactivity results when stable nuclei are subjected to bombardment by particles.
- b. If the energy of the particle is of the proper value, the particle will combine with the nucleus. The resulting nucleus, if unstable, will be radioactive.

D. Isotopes

1. Kinds of isotopes.⁵

- a. Stable.
- b. Radioactive.

2. Production of isotopes.⁶

- a. Cyclotron.
- b. Reactor.

² Charles E. Dull; H. Clark Metcalfe; and John E. Williams, *Modern Chemistry* (New York: Henry Holt and Company, 1958), 556-557.

³ *Nuclear Science Teaching Aids and Activities* (Washington: Department of Health, Education, and Welfare, June 1960), pp. 31-35. (reprint)

⁴ Samuel Glasstone, *Sourcebook of Atomic Energy* (second edition; Princeton, New Jersey: D. Van Nostrand Company, Inc., 1958), 273-312.

⁵ *Ibid.*, pp. 197-236.

⁶ *Ibid.*, pp. 237-272; 449-481.

¹ G. M. Bradbury and others, *Chemistry and You* (Chicago: Lyons and Carnahan, 1957), 331-367.

3. Characteristics of useful isotopes.
 - a. Half-life of 30-minutes or longer.
 - b. Sufficiently energetic radiation to insure easy detection.
4. Methods of separating radioisotopes.
 - a. Gaseous diffusion method.
 - b. Electromagnetic method.
 - c. Thermal diffusion.
 - d. Centrifugal method.
 - e. Electrolysis.
 - f. Chemical exchange (ion).
 - g. Distillation.
5. Physical applications.¹
 - a. Tracer techniques.
 - b. Absorption studies.
6. Chemical applications.
 - a. Exchange reactions.
 - b. Szilard-Chalmers type reaction.
- E. Nuclear particles requiring detection and their properties.
 1. Alpha particles.
 2. Fission fragments.
 3. Beta particles.
 4. Gamma rays.
 5. Neutrons.
- F. Survey of detection methods. ("Detection" intended to include not only the presence of nuclear radiation but also the measurement of the amount of energy and related properties.)

Each pupil can construct his own electroscope for detection of radioactivity.

Continuous cloud chamber: Construct a continuous chamber as outlined in Frank Starr's thesis. A simpler version may be constructed using a small peanut butter jar.

Film: *Unlocking the Atom*. United World Films, Inc., 20 min., 16mm, sd., b&w. Available on free loan from Public Information Service, U.S. A.E.C., Washington 25, D.C. To acquaint pupils with principles which govern the atom and its use.

1. Types of detectors.
 - a. Gas-filled detectors.
 - (1) Ionization chambers.

Bring a flame (match) near a charged electroscope. The electroscope will lose its charge.

¹ "Absorption—plants and Absorption—animals," *Nuclear Science Teaching Aids and Activities* (Washington: Department of Health, Education and Welfare, June 1960), 31, 33. (reprint)

Bring a radioactive material (as supplied with a classmaster) near a charged electroscope; the electroscope will be discharged.

Perform experiment in Leybold manual. Attach the Leybold ionization chamber to the Leybold-Wulf electroscope. The air inside the ionization chamber is nonconductive until a radioactive material is introduced inside the chamber. If a voltage is applied across the ionization chamber, a current can be detected by using the Wulf electroscope as a sensitive ammeter. The saturation current can also be shown.

- (2) Proportional counters.
- (3) Geiger-Muller tubes.

Explain how a Geiger-Muller tube operates. Demonstrate the characteristic curve and operating voltage of a Geiger-Muller tube. Determine the background count.

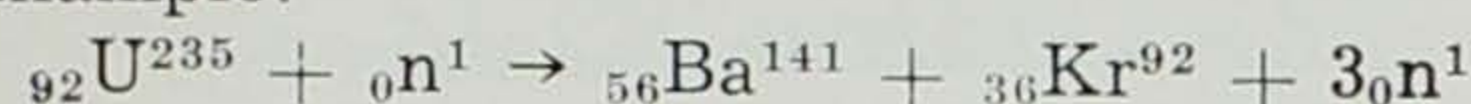
- b. Scintillation detectors.
- c. Cloud chambers.
- d. Nuclear track plates.
- e. Crystal counters.
- f. Cerenkov counters.
- g. Chemical detectors.
- h. Calorimetric methods.
- i. Neutron detectors (various types).
Types c. to i. are specialized.
2. Radiation detection systems.
 - a. Particle counting.
 - b. Pulse height analysis.
 - c. Coincidence measurements.
 - d. Mean level detection systems.
 - e. Special techniques.
 - (1) Nuclear emulsion.
 - (2) Cloud chamber.
3. Statistics of counting units of measurement.
 - a. Curie—subdivisions.
 - b. Roentgen.
 - c. Rad.
 - d. RBE.
 - e. REP.

G. Fission reactions.

1. Meaning of fission.²

² Glasstone, *op. cit.*, pp. 385-412.

Example:



2. The fission process.¹

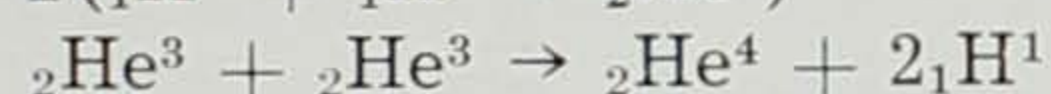
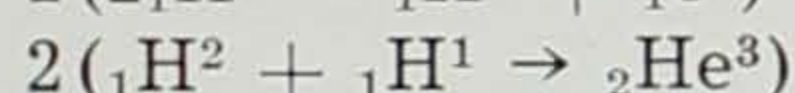
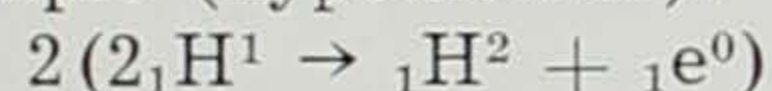
Using U^{235} as an example, discuss fission reactions in detail, following readings in outline.

- The production of from 1 to 4 neutrons during each fission is needed to promote the chain reaction.
 - Fission fragments are radioactive isotopes ranging in atomic number from about 30 to 64 in atomic number.
 - The conversion of about 0.2 atomic mass units into energy takes place during each fission reaction.
 - Most of the energy released appears as kinetic energy of the fission fragments.
3. Controlled fission.²
Application of controlled fission in reactors to produce energy and new isotopes.

Use the diagram of a reactor that is used in the book, *Modern Chemistry*, and discuss fundamental principles of a reactor involving controlled fission.³

H. Fusion reactions.

Example (hypothetical):



1. The fusion process as distinguished from the fission process.

Give examples of various hypothetical fusion equations. Contrast these with other types of nuclear reactions involving fission or disintegration.

2. Fusion as the source of energy of the sun.⁴

Discuss fusion of hydrogen to helium as the postulated source of the energy of the sun.

3. Fusion processes on earth.

Account for the difficulties achieving fusion on the earth.

Show possible fusion equations taking place in hydrogen bomb explosions.

Discuss the possibilities of developing controlled hydrogen fusion reactions as a source of energy for peaceful use.

- Hydrogen bomb.
- Controlled fusion—a future source of energy.

I. Applications of nuclear chemistry.

1. Tracers—a tracer is a radioactive isotope with a characteristic radiation by which it can be distinguished and the absolute amount present measured in the presence of similar non-radioactive species.⁵

a. Tracer techniques in biology.⁶

- To illustrate detection of radioactivity in biological materials with film.
- Method to show plant absorption and distribution of minerals through excised stems.
- Isotopes may be used in plant translocation.
- Assimilation.
- Foliar feeding.
- Plant injection.
- Selective uptake.
- Sampling biological materials.

b. Tracer techniques in chemistry.⁷

- Equilibrium—dissolved or undissolved salts in a solution.⁸
 - To show interchange of ions.
 - To show initial rate of exchange can be increased.
- Solubility—determining of the solubility of a salt in water.⁹
- Concentration determinations by isotope dilution.¹⁰
- Isotope separation.¹¹

⁵ Willard F. Libby, "Developments in the Peacetime Uses of Atomic Energy," *Journal of Chemical Education*, XXXVI, 627-633.

⁶ Frank Starr, "Radioactivity, Demonstrations, Experiments and Techniques Using Radioisotopes in Science Teaching," unpublished Master's thesis (Cedar Falls: State College of Iowa).

⁷ Friedlander, Gerhard, and Kennedy, *Nuclear and Radiochemistry* (New York: John Wiley and Sons, Inc., 1955).

⁸ Starr, *op. cit.*, p. 54.

⁹ Starr, *op. cit.*, p. 55.

¹⁰ Starr, *op. cit.*, p. 56.

¹¹ Starr, *op. cit.*, p. 57.

¹ *Ibid.*

² Glasstone, *op. cit.*, pp. 413-436.

³ Charles E. Dull; H. Clark Metcalfe; and John E. Williams, *Modern Chemistry* (New York: Henry Holt and Company, 1958), 567-568.

⁴ Glasstone, *op. cit.*, pp. 440-448; Paul Frey, *College Chemistry* (Englewood Cliffs, New Jersey: Prentice-Hall, Inc., 1960), 698.

- (5) Radioassay of ores.
- (6) Radioisotopes in chromatography.¹
- (7) Separation and ion exchange.
- (8) Tracers can be used for determining the structure of compounds.
- (9) Tracers can be used for the study of the mechanisms of chemical reactions.

Ester hydrolysis.

c. Tracer techniques in medicine.

- (1) For studies of blood circulation.
- (2) For location of tumors.
- (3) Other.

d. Tracers in agriculture.

Film: *Atoms in Agriculture*, 16mm sound, b&w, E.B.F. Free from Public Information Service, A.E.C., Washington 25, D.C.

e. Physics.

- (1) Research on the various phases of radioactivity.
- (2) Investigation of wear.

f. Industry.

- (1) Non-destructive testing and control of thickness and quality.
- (2) Searching for leaks.
- (3) Treatment of products to improve properties.
- (4) Sterilization.
- (5) Source of power.

* *Rules for Handling Radioactive Materials*

1. Use only those quantities of radioactive materials for which no licensing or special regulations are required.
2. Be sure you know what nuclear reaction is producing the radiation with which you are working.
3. Know exactly what quantity of unstable nuclei for which you are responsible.
4. Acquaint yourself with the characteristics of the radiation emitted by whatever unstable nuclei with which you are working.
5. Before obtaining the necessary radioactive materials, be sure all of your exercises are sufficiently well planned to allow you to use the short-lived materials before they undergo extensive decay.
6. From the moment your radioactive materials arrive, be sure your storage facilities are such that you will know precisely where each bit of it is until you are through working with it and it has been flushed down the drain or otherwise disposed of.

—from *Nuclear Science Teaching Aids and Activities*.
#MP14, p. 31.
by Office of Civil Defense, Washington, D.C.

¹ Starr, *op. cit.*, p. 58.

SUPPLEMENTARY LECTURES, TESTS AND SUGGESTED LIST OF CHEMICALS

BEHAVIOR OF GAS MOLECULES

In most texts the behavior of gases is limited to a discussion of their general nature and to a presentation of Boyle's and Charles' laws solely from the standpoint of reduction of gas volumes to standard conditions of temperature and pressure. While this is adequate for laboratory purposes and as a background for the discussion of molecular weights, it leaves much to be desired. It is the opinion of the committee that a more complete coverage of the molecular-kinetic behavior of gases should be included in a really good chemistry course.

The following presentation is intended to supplement, not displace, the material in the regular text. It may be used as a whole or in part for a lecture, or it may be mimeographed, distributed to the class, and used for study and discussion.

Skin diving, or more exactly SCUBA diving, and swimming have recently become popular sports. We read of the exciting underwater explorations of these swimmers and see and hear over television thrilling dramas taking place almost entirely on the bottom of the sea. Few think, however, that each and every one of us lives out his life at the bottom of a sea of air much deeper and even more interesting in its properties than the seas of water.

We are so familiar with living conditions in our sea of air that it seems not half as mysterious as the ocean. But each has its currents, its change of pressure with depth, its plants and animals. The occasional deep disturbances in the water seas are as nothing compared to the tremendous storms taking place constantly in some portion or another of our sea of air. Even though we are familiar with the bottom of our sea of air, and perhaps a layer a few miles up from the bottom, it is only recently that exploration of the upper air has become possible. This is comparable to the comparatively recent fragmentary groping in the greater depths of our water seas.

There is still much to be learned of both these fluids in or near which we live. A careful study of the facts already known about both our seas is necessary before we can explore them further for new wonders.

Our sea of air is composed of gas or rather of a mixture of gases. As you already know, nitrogen and oxygen are present in the greatest abundance,

together with carbon dioxide, water vapor, and smaller percentages of a half dozen rarer gases. Its composition is remarkably uniform in our layer near the bottom, but this changes as we go upward so that the upper layers are richer in the lighter components, even hydrogen being present at extreme altitudes.

Of the three states of matter, gas, liquid, and solid, the gaseous state is by far the simplest. Nevertheless, when studied in detail, even this becomes terribly complex. In fact, without considerable skill in higher mathematics, the rigorous study of gas behavior called Kinetic Theory is impossible. It is possible, however, to learn a great deal about gases with no more than simple algebra.

For example, when we blow up a toy balloon with air, what happens? Of course the balloon becomes larger and larger and more firm to the touch as we force more air into it. We know from earlier study that the air is composed of minute molecules of nitrogen and oxygen. We also know that we cannot be completely filling the balloon with molecules as we would a bag with marbles. In the first place, the "filled" balloon is not heavy enough. Oxygen molecules are fairly heavy, and enough oxygen and nitrogen molecules packed solid in a ten centimeter or four-inch diameter balloon would weigh about 580 grams or a little over a pound. Furthermore, the balloon is still elastic. Not only do the rubber walls stretch; but if we squeeze the balloon as a whole, we can force it into a smaller volume. Experience with tire pumps and similar vessels enclosing gases shows that gases are mostly space.

Calculation of the weight of gas in our four-inch balloon or even in a tightly blown up tire indicates little actual material present in our gas for the space it occupies. Even though there is little weight of material present, the molecules are so small that there are billions of them present.

Pressure is force exerted per unit area of surface. The only way we can account for the production of so much pressure on the inside surface of our balloon is by imagining the molecules to be moving at high speeds and constantly banging into the inside walls of their container. Each time a molecule collides with the wall of the balloon, it exerts a minute force upon the wall; and billions of these tiny impacts on a square centimeter of

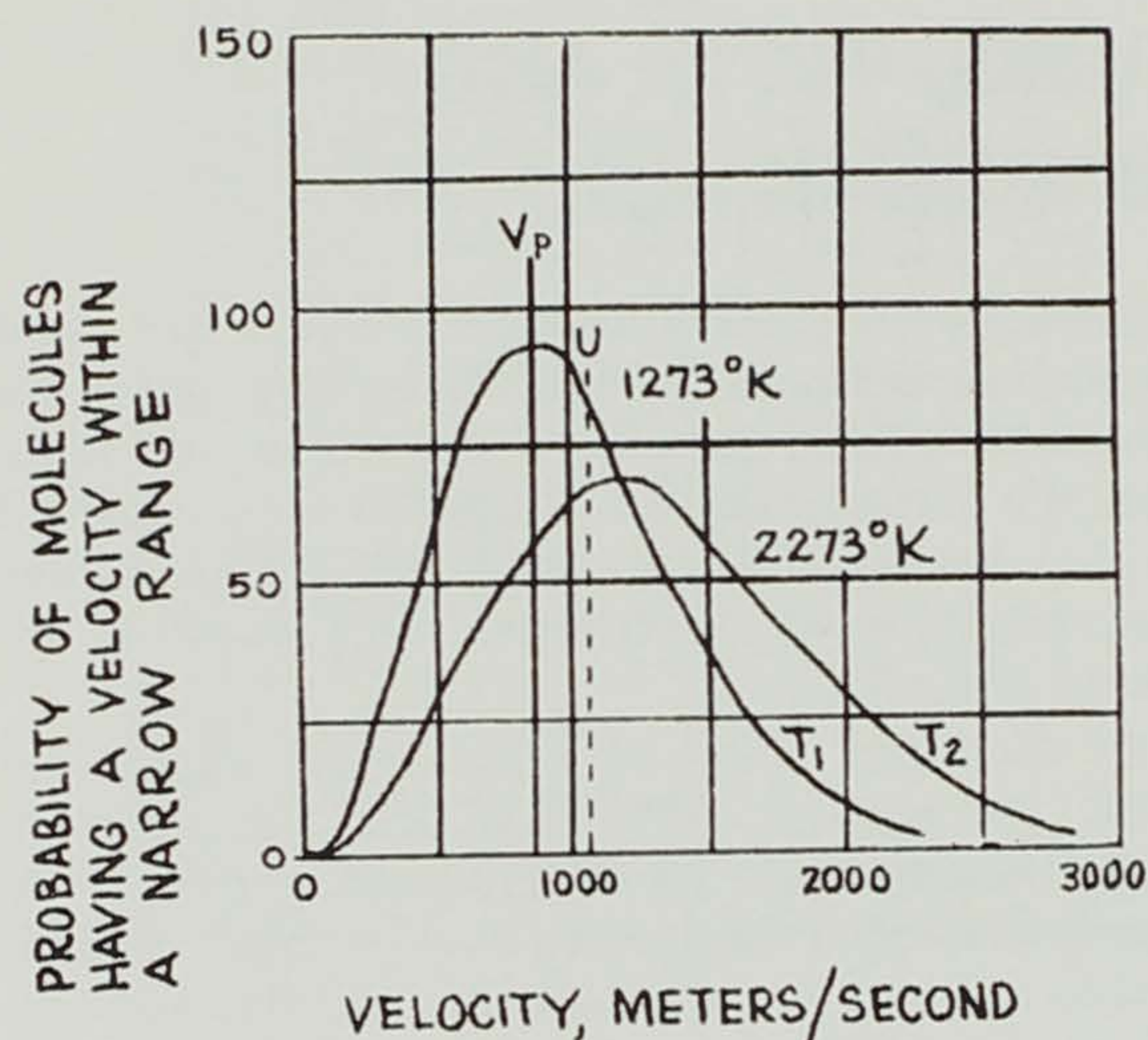


Figure 18

surface combine to give the total force which we call pressure.

The actual number of these collisions per second, even at room temperature and at 74 cm (the pressure within our balloon), is tremendous. It amounts to about 4×10^{23} impacts per square centimeter of surface per second. Not only do the molecules collide with the walls, but they are also bound to collide with each other, so that the total number of molecular collisions may run as high as 1.7×10^{34} per second, or 17,000,000,000,000,000,000,000,000,000,000.

If we heat the balloon and its gas, the balloon becomes larger and is stretched more tightly. In other words, the pressure of the gas inside has increased. Since no more gas molecules have been added to the balloon, each molecule must be "working harder"; that is, the typical molecule must either be hitting the walls harder, or more often, or both. Actually, the latter is the case.

To understand this, it is necessary to imagine the situation within the gas. Here we have a large number of molecules flying around completely at random like a swarm of angry bees. At a given instant a very few will be standing still. Others will be moving with extremely high speeds, even higher than that of the bullets from a high-powered rifle. Most will be moving at intermediate speeds or perhaps we should say *velocities* to be scientifically correct. By far the greatest number will have what is known as the most probable velocity V_p indicated by the peak of the curve, FIGURE 18. The geometric average velocity, U , which

actually determines the pressure that the gas molecules exert is equal to the square root of the sum of the squares of all the velocities:

$$U = \sqrt{\frac{1}{N} \sum C^2}$$

This *root mean square* velocity is a little higher or 1.224 times the most probable velocity and is indicated by the vertical dotted line on the curve T_1 . Curve T_2 represents the behavior at a higher temperature.

DEMONSTRATION OF MOLECULAR MOTION

Molecular motion may be simulated in several ways. The simplest is the method of Minneart which consists merely of enclosing a number of steel balls in a rectangular wood frame with a masonite bottom. The frame, which should be about one foot square, is laid on a low table in full view of the class and given a slight rotary motion with the hands. Increasing the rate of motion represents an increase in temperature. The greater number of impacts per second which results gives a greater "pressure" for the constant "volume" of the frame. Any diameter balls from the size of BBs to $\frac{3}{8}$ inch ball bearings may be used, but at the start of the demonstration they should be of uniform size. After the first showing, it is instructive to introduce a few larger balls with the smaller ones to illustrate the point that heavy molecules move slower, while light molecules move faster; but at the same temperature the energy of the impacts with the walls is the same so that the pressure developed in a given container by a given fraction of a mole is the same for different gases. That is:

$$P_1 V_1 = P_2 V_2 = nRT$$

$$\frac{2}{3} \cdot \frac{1}{2} M_1 U_1^2 = \frac{2}{3} \cdot \frac{1}{2} M_2 U_2^2 = nRT$$

$$\text{KE}_1 \qquad \qquad \text{KE}_2$$

Where $\frac{1}{2} MU^2 = \text{Kinetic Energy (at constant } T)$

If an overhead slide projector is available, a more sophisticated demonstration, visible to a larger class, can be arranged. See FIGURE 19. From $\frac{3}{16}$ inch transparent polystyrene sheet, obtainable from any radio supply house, saw a rectangular piece of plastic a little larger than the optical opening in the stage of the projector. For projectors taking standard sized slides, the piece of plastic should be cut approximately $3\frac{3}{4} \times 4\frac{1}{4}$ inches for such a machine. Cut also several strips $\frac{1}{4}$ inch wide and long enough to form a frame when cemented around the edge on top of the larger piece of plastic. Polystyrene cement can be obtained

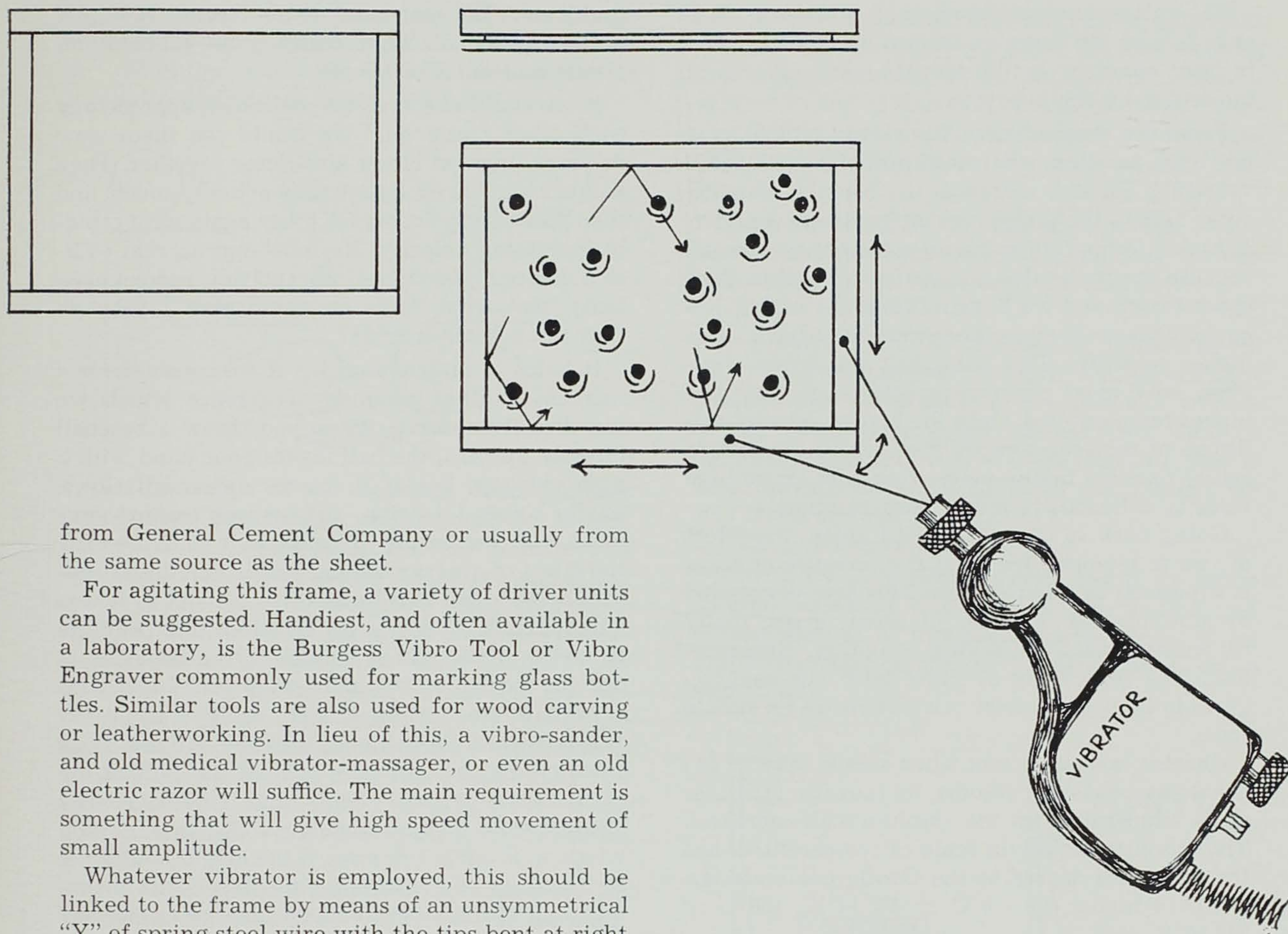


Figure 19

from General Cement Company or usually from the same source as the sheet.

For agitating this frame, a variety of driver units can be suggested. Handiest, and often available in a laboratory, is the Burgess Vibro Tool or Vibro Engraver commonly used for marking glass bottles. Similar tools are also used for wood carving or leatherworking. In lieu of this, a vibro-sander, and old medical vibrator-massager, or even an old electric razor will suffice. The main requirement is something that will give high speed movement of small amplitude.

Whatever vibrator is employed, this should be linked to the frame by means of an unsymmetrical "Y" of spring steel wire with the tips bent at right angles at the ends of the "Y" and pressed into holes bored near one corner of the plastic frame. See diagram. For use with the Burgess tool, the tail of the "Y" is soldered into a short piece of $\frac{1}{4}$ inch copper tubing and this held in the chuck of the tool. With other vibrators, the attachment will have to be left to the ingenuity of the demonstrator.

In operation, small balls are placed in the frame and agitated by the vibrator mechanism in a manner similar to that used with the larger model of Minneart described above. The size and weight of the balls depends somewhat on the degree of agitation employed and must be determined by trial and error. No. 4 lead shot usually works quite well. BBs are a little large for the size of the frame suggested. For light vibration, radish seed is very effective, and rape seed may be used for larger balls. The shadows of the rapidly moving balls are projected on the screen and with sharp focus and under ideal conditions appear to be moving in three

dimensions as indeed they are to some extent. In fact, there is some tendency for the "molecules" to jump out of the frame and "diffuse" onto the floor if too great agitation is employed.

If the amplitude of vibration can be controlled by means of an adjustment on the vibrator or by use of a rheostat in the powerline, it is very effective to start with the "molecules" at rest or "absolute zero" with them arranged into an almost perfect "crystal" lattice. Then with the frame tilted slightly toward the bottom, agitation is started and the "crystal" melted to a "liquid." Some "molecules" escaping from the surface into the gas phase represent evaporation. As the agitation is increased, the boiling point is reached and all the molecules "evaporate" into the gas phase.

If a temporary partition made with a small gate is wedged across the center of the frame, Graham's law of diffusion may be illustrated by using two different weights of balls.

The entire demonstration, once prepared, takes only five or ten minutes from a class period and is very effective in illustrating many phases of molecular-kinetic behavior.

From this we can see that a rise in temperature not only increases the most probable and r.m.s. velocities but also increases the fraction of molecules having the higher velocities. Thus, we have, at higher temperatures, more collisions per second because the molecules are getting around in their space faster, and we have more of the higher impact collisions because more molecules have these higher velocities. This latter fact also tends to explain why most chemical reactions go faster at higher temperatures. Only the higher velocity collisions of molecules capable of reacting result in actual reaction taking place; therefore, more high velocity collisions produce more reaction.

Going back to our gas system again, the effect of rise in temperature upon the pressure of a gas is expressed by Charles' law. This law, discovered by studying the behavior of actual gases, really represents exactly the ideal situation. Nevertheless, it is sufficiently accurate with our common gases to be used in most calculations in its simple form.

Charles' law states that when a fixed mass of gas is held at a constant volume, its pressure increases as its temperature on the absolute scale increases. The absolute or Kelvin scale of temperatures has the same size degree as the Centigrade scale but is 273.14 higher, i.e.: $0^{\circ}\text{C} = 273.14^{\circ}\text{K}$; $100^{\circ}\text{C} = 373.14^{\circ}\text{K}$, etc.

Since Charles' law indicates a direct proportion between pressure and temperature, we may write $P \propto T$ or $P = kt$ (V constant). (1)

This simple law is actually quite fundamental, for as you study further in physics or in chemistry, you will find that the temperature of a body is actually a measure of the intensity of the molecular motion in the body. The Kelvin scale of temperature is really based upon the ideal behavior of gases, so it is not surprising that a close relationship exists between gas pressure and temperature.

Now supposing we take a hot gas and enclose it in a perfectly insulated container. What happens? On the basis of the common behavior of colliding bodies, we would expect the gas to cool down and its pressure to drop. Marbles rolling around in a flat pan strike each other a few times and soon come to rest. Even billiard balls, which are very elastic, lose some energy when they strike each other or the walls of their container. After several collisions they slow down, hit less hard, and even-

tually stop. But molecules do not lose energy upon normal collision! What really takes place in an elastic molecular collision?

If we could observe two molecules approaching each other "head on," we would see them slow down as they get closer and closer together. They would come to rest before they actually touch, and then move away from each other again with rapidly increasing velocity. No touching, no real collision, hence no frictional loss, and thus perfect elasticity. But what "puts on the brakes"? Why do they not actually collide?

It helps to understand this if we consider for a moment another common occurrence which we can directly observe. When you throw a baseball directly upward, the ball leaves your hand with a high velocity; but as it travels upward it slows, finally comes to a stop, and returns toward your hand with increasing velocity as it falls. Barring the effect of air friction, the ball is moving as rapidly by the time it returns to your hand as it was when it left. There is no collision here, but the behavior of the ball is similar to the behavior of our gas molecule. The ball leaves your hand with a certain amount of kinetic energy, $E_{\text{kinetic}} = 1/2MU^2$ where M is the mass of the ball and U its velocity. As the ball rises against the pull of the earth's gravitational field, this kinetic energy changes into potential energy. $E_{\text{potential}} = mhg$ where m is again the mass, h the height, and g the acceleration of gravity. At the top of its movement, all of the kinetic energy (energy of motion) has changed into potential energy (energy of position) so the ball stops. On the way back down the reverse process takes place.

A similar thing happens to our approaching molecules which we must remember are electrical in nature. Each is surrounded by an electrical field due to the many electrons in its structure. We do not know as much as we would like about the exact nature of electrical fields anymore than we do about the gravitational fields of our earth and of other bodies. We do know, however, that like electrical fields repel each other. Since the electrons in both our approaching atoms are negative charges, their fields repel each other* as the atoms come closer together.

Normal fields of molecules are roughly spherical, or in some cases, egg shaped. We must think

* There are also weak attractive forces between gas molecules which are felt at a distance. These are due to other effects which we need not consider here as they are overshadowed by the strong forces of repulsion acting at the distances we are now considering.

of these normal fields becoming more and more flattened on the surfaces opposed to each other as the molecules approach. Perhaps the picture is similar to that shown below (FIGURE 20).

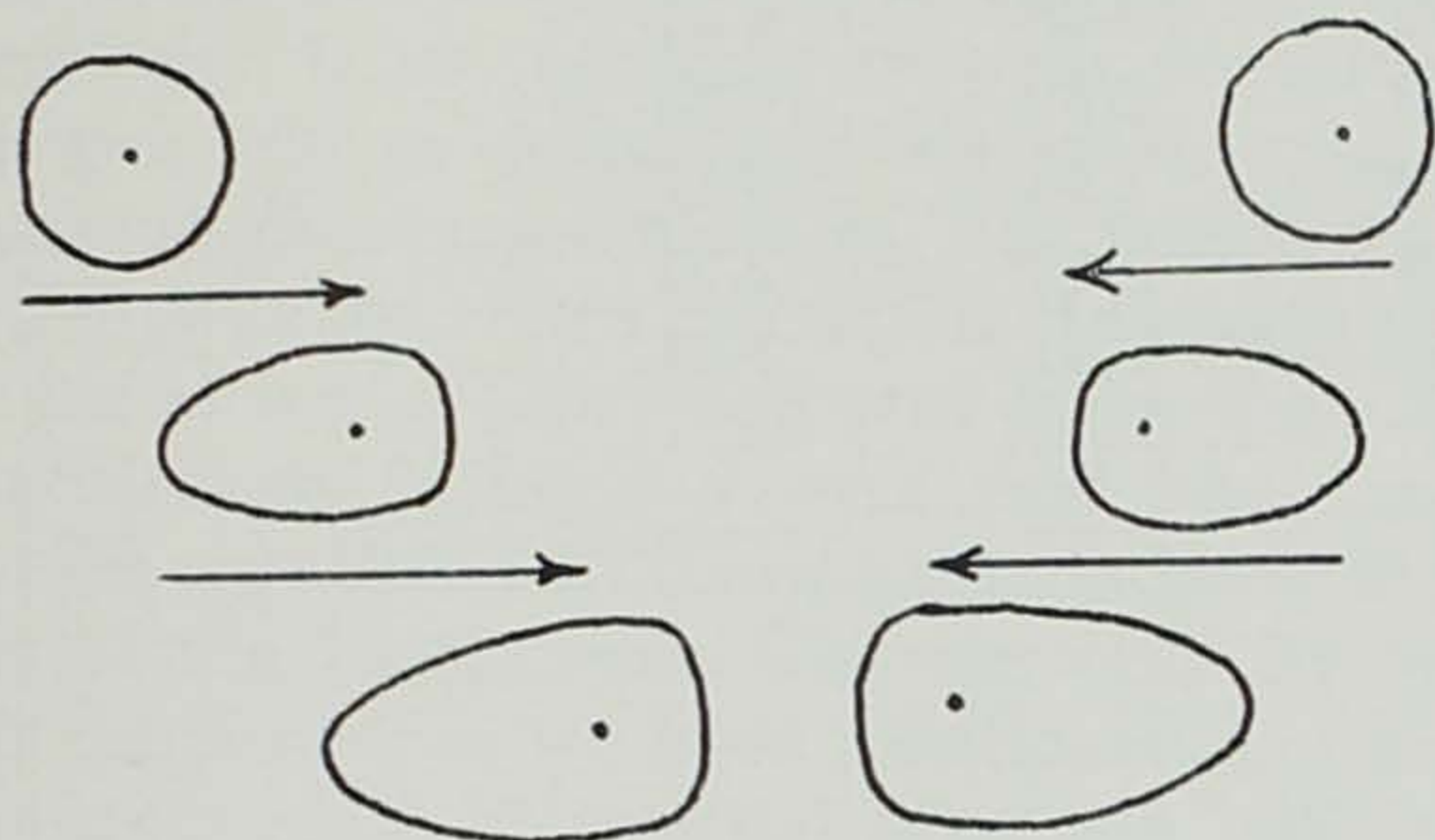


Figure 20

This distortion of the electrical fields requires energy and this energy comes from the kinetic energy of the moving molecules. As the molecules approach closer and closer together, more and more of the kinetic energy of their motion is changed into potential energy of electrical distortion or strain. At some point, when the molecules are very close, all the energy of motion has become potential so the molecules stop. This is the "collision" and the distance of closest approach is considered to be the effective diameter of the molecule. That is to say, it is a measure of the space the molecule occupies to the exclusion of all other molecules.

It is interesting to know that this apparent diameter becomes slightly smaller as the temperature of the gas is raised. Can you explain this? (Note: Good topic for class discussion.)

The statement was made previously that the temperature of a body is a measure of the intensity of the molecular motion in the body. This is indeed true; and after much detailed study of the behavior of gases, scientists have arrived at an equation for ideal gases which directly relates the kinetic energy of linear motion of the molecules to their absolute temperature.

$$E_k = 1/2 MU^2 = 3/2 RT \quad (2)$$

Where R is a constant for ideal gases and one which is applicable to all actual gases within reasonable limits of error as long as we do not go to extremes of temperature and pressure. If the above is true, we can also write:

$$2/3 \cdot 1/2 MU^2 = 1/3 MU^2 = RT \quad (3)$$

In this and the above equation, the M is not the mass of a single molecule, but rather the mass of

6.02×10^{23} molecules that we call a molecular weight. This molecular weight of gas, representing such a large number of molecules, will occupy a volume V , and at this volume will produce a pressure P due to the millions of impacts of these rapidly moving molecules on the walls of their container.

Since this pressure-volume effect is due to the kinetic energy of the molecules, we may include these terms in the above equation and write:

$$PV = 1/3 MU^2 = RT \quad (4)$$

This accords with the findings of Robert Boyle who discovered as early as 1662 that, at constant temperature, the product of pressure and volume of a specified mass of gas was constant. Thus in the above equation, R is a fixed constant; and if we arbitrarily hold T constant for the moment, we may write our equation:

$$PV = RT = \text{constant} \quad (5)$$

Now if $PV = k$ (Boyle's law) and if we insert different values of V in the equation, we find that P decreases as V increases. Furthermore, if we graph this relation we get a figure known as an hyperbola:

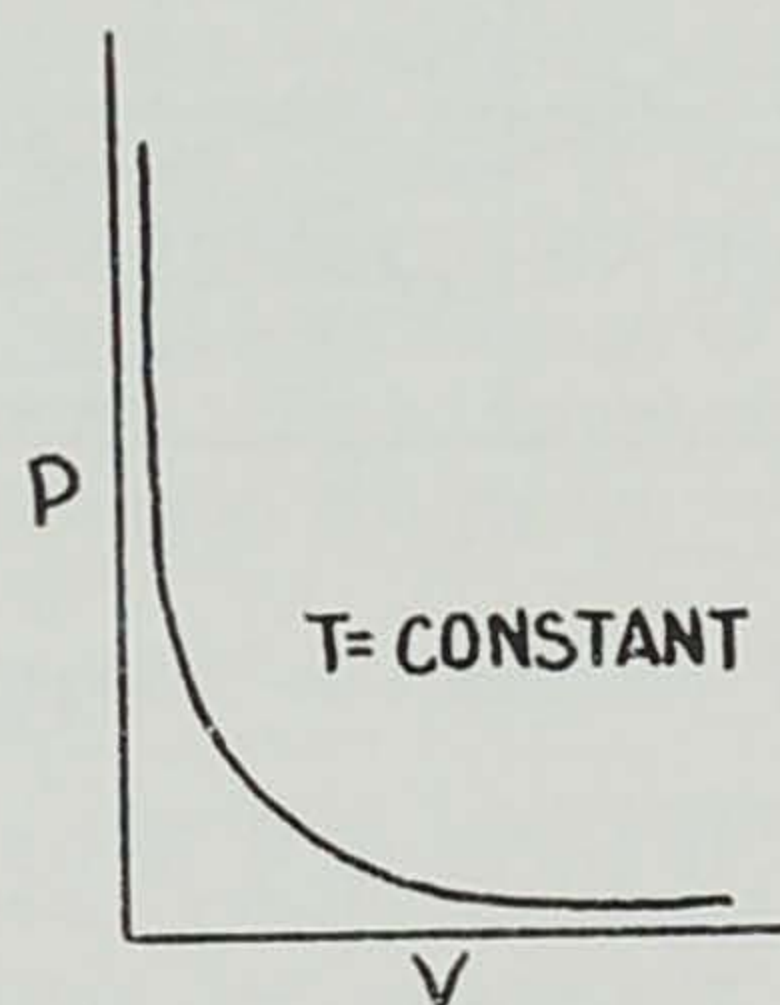


Figure 21

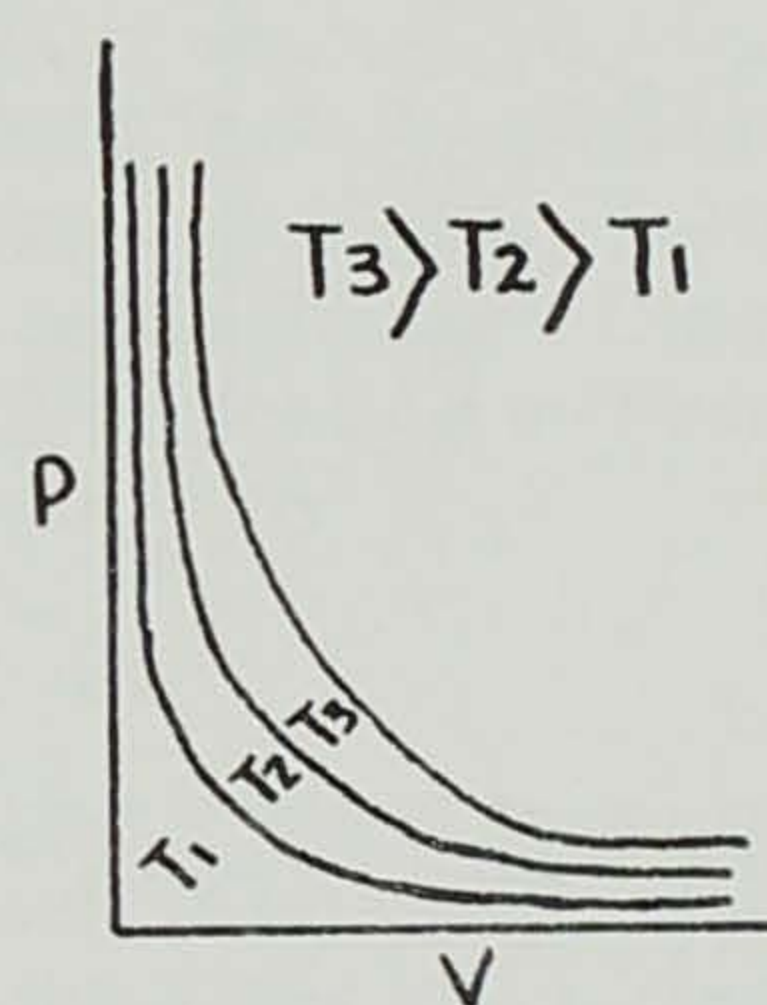


Figure 22

and since this is at a constant temperature, we call the curves an isotherm for an ideal gas at T . We have already learned that at any constant volume the pressure increases as the temperature increases, so we can graph a "family" of isotherms, each at a little higher temperature and so depict graphically the entire behavior of an ideal gas. The actual points shown are calculated from the equation as follows:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} = \frac{P_3 V_3}{T_3} = 1 R = 0.0821 \text{ liter atmospheres per degree.}$$

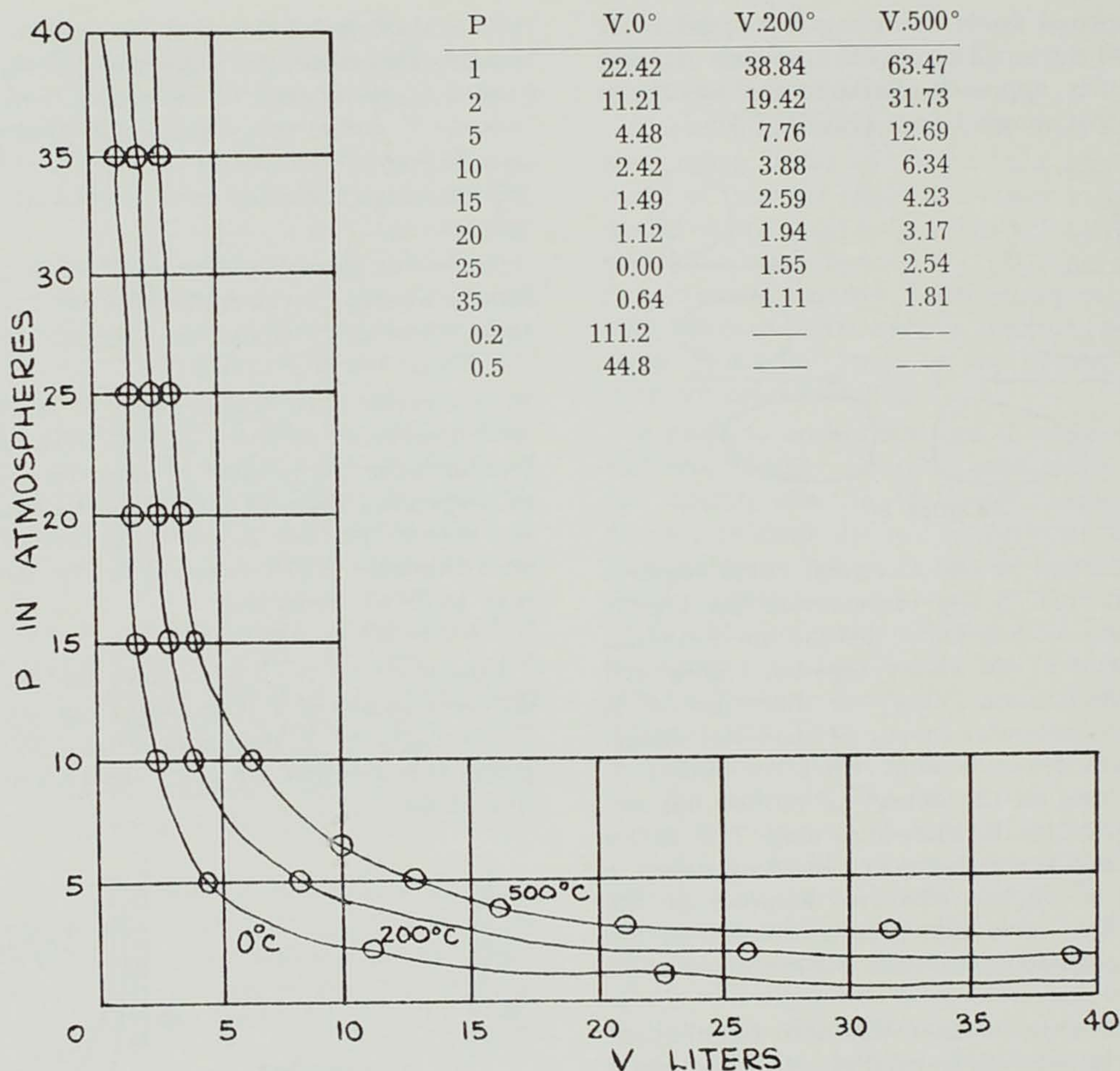


Figure 23

This is for one mole of gas. If we have $\frac{1}{2}$ or 0.5 mole of gas, the values of $\frac{PV}{T} = 0.5R$. For any

number of moles or decimal fraction of a mole, n , the equation becomes:

$$PV = nRT \text{ or } \frac{PV}{T} = nR \quad (6)$$

This is one of the most useful equations in physics and chemistry and with it and equation (3) most problems relating to gases may be solved.

As pointed out several times, the equations shown are for an ideal gas.

Our picture of an ideal gas is one in which the molecules are points, not taking up any space, and in which they have no attraction for each other.

Obviously no actual gas behaves in this way. The molecules behave as if they had a fairly definite diameter. They also show a greater or lesser attraction for each other, depending on the gas; and this attraction is felt over a distance of several molecular diameters. It is this molecular attraction which results in the liquefaction of gases when they are cooled and compressed and which also accounts for the stability of liquids and many solids at ordinary temperatures. No actual gas is ideal; but at low and moderate pressures, the results of calculations using the ideal gas equation are sufficiently accurate for most purposes. For more accurate results, we must use equations such as that of van der Waals which introduces correction factors into the ideal equation for the size of the molecules and the attractions between them.

So we have learned some of the peculiarities of gases and the main equations describing their behavior.

However, this still leaves our sea of air largely unexplored. If you continue on your scientific career, you will go deeper and deeper into the systematic study of gas molecules. You will find some as strange as the weird fishes pulled from the deep troughs of our oceans. You will learn that many molecules go flying end over end like thrown boomerangs. You will learn about others, in fact most of the larger ones, that not only fly about their containers as we have mentioned, but are in a constant state of internal vibration with their atoms lolloping around in a most surprising fashion as if loosely tied together with a lot of springs. It has been said that science is our last unexplored horizon.

A SIMPLE pH METER—PUPIL PROJECT

The antimony electrode is sensitive to changes in pH and furnishes a rugged, easily-made electrode for approximate pH measurement. The saturated calomel electrode is also easily reproduced and acts as a reference electrode. The voltage of the cell resulting from the combination of these two electrodes dipping into the solution to be measured can be found with reasonable accuracy by Hildebrand's potentiometer. This can be easily assembled from components present in most laboratories or purchasable in inexpensive form from radio supply houses.

Materials needed. Small amount of antimony metal of reasonable purity, about $\frac{3}{8}$ inch of fine platinum wire, glass tubing, small globule of mercury, 5 g mercurous chloride, C.P., 50 g potassium chloride C.P., sealing wax, voltmeter reading 0-1 v or milliammeter 0-1 ma or 0-10 ma (in which case a suitable multiplier resistor is needed), simple galvanometer, 100 or 200 ohm radio potentiometer (wire wound) with knob, s.p.s.t. toggle switch, push button, No. 6 dry cell.

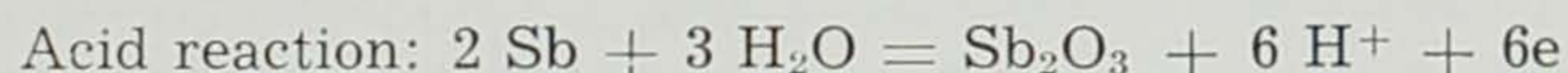
Theory. The antimony-antimony oxide electrode is theoretically reversible, and the chemical reaction involves either the H^+ or OH^- ion according to the acidity of the solution. The chemical equation and form of the Nernst equation for calculating the E.M.F. of the electrode half-cell are given for each case.

The calomel electrode is well known and adequately described in most texts.

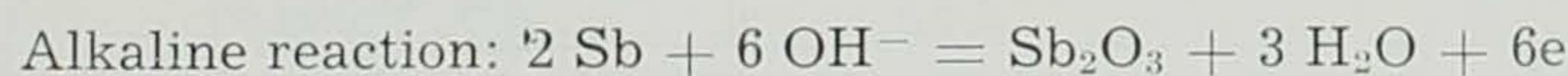
The cell composed of these two electrodes and the solution to be measured, which constitutes the electrolyte, has a moderately high resistance. For this reason, the true reversible voltage can only be measured by opposing the voltage of the cell by that from a potentiometer. The circuit of Joel Hildebrand furnishes a true potentiometer, simple and satisfactory for this purpose, although its accuracy depends upon that of the voltmeter used. In the circuit illustrated, the partial voltage of the working cell "W" as obtained from the potentiometer resistance is measured by means of a voltmeter and balanced against the voltage of the experimental cell. When balance is obtained, the galvanometer will read zero. The energy to operate the voltmeter comes from the working cell and no current is drawn from the experimental cell at the condition of balance. Thus the reversible or true voltage of this cell is obtained. Directions: (To be supplemented by reading and consultation with your instructor.)

Making of Antimony Electrode

Carefully melt some antimony metal in a crucible (avoid breathing fumes). Skim off the dross from the surface and cast the metal into a short stick by pouring it into the end of a small piece of 5 mm pyrex glass tubing. The tubing should be stoppered by a piece of asbestos in the lower end and well supported over a tray so that spilled metal will not burn the table top. When the metal has cooled, break off the glass and scrape the stick of metal clean. Solder an 18-inch length of stranded



$$\begin{aligned} E_{\text{Sb}} &= E^\circ_{\text{Sb}} - RT \ln a_{\text{H}^+}^6 = \frac{-0.0591}{6} \log a_{\text{H}^+}^6 \\ &= -0.1445 + 0.0591 (\text{pH}) \quad \text{At } 25^\circ\text{C} \end{aligned}$$



$$E_{\text{Sb}} = E^\circ - \frac{RT}{6f} \ln \frac{1}{a_{\text{OH}^-}^6} = 0.6834 - 0.0591 \log \frac{1}{a_{\text{OH}^-}^6} \quad \text{At } 25^\circ\text{C}$$

$$E (\text{saturated calomel}) = 0.2458 \text{ volts.}$$

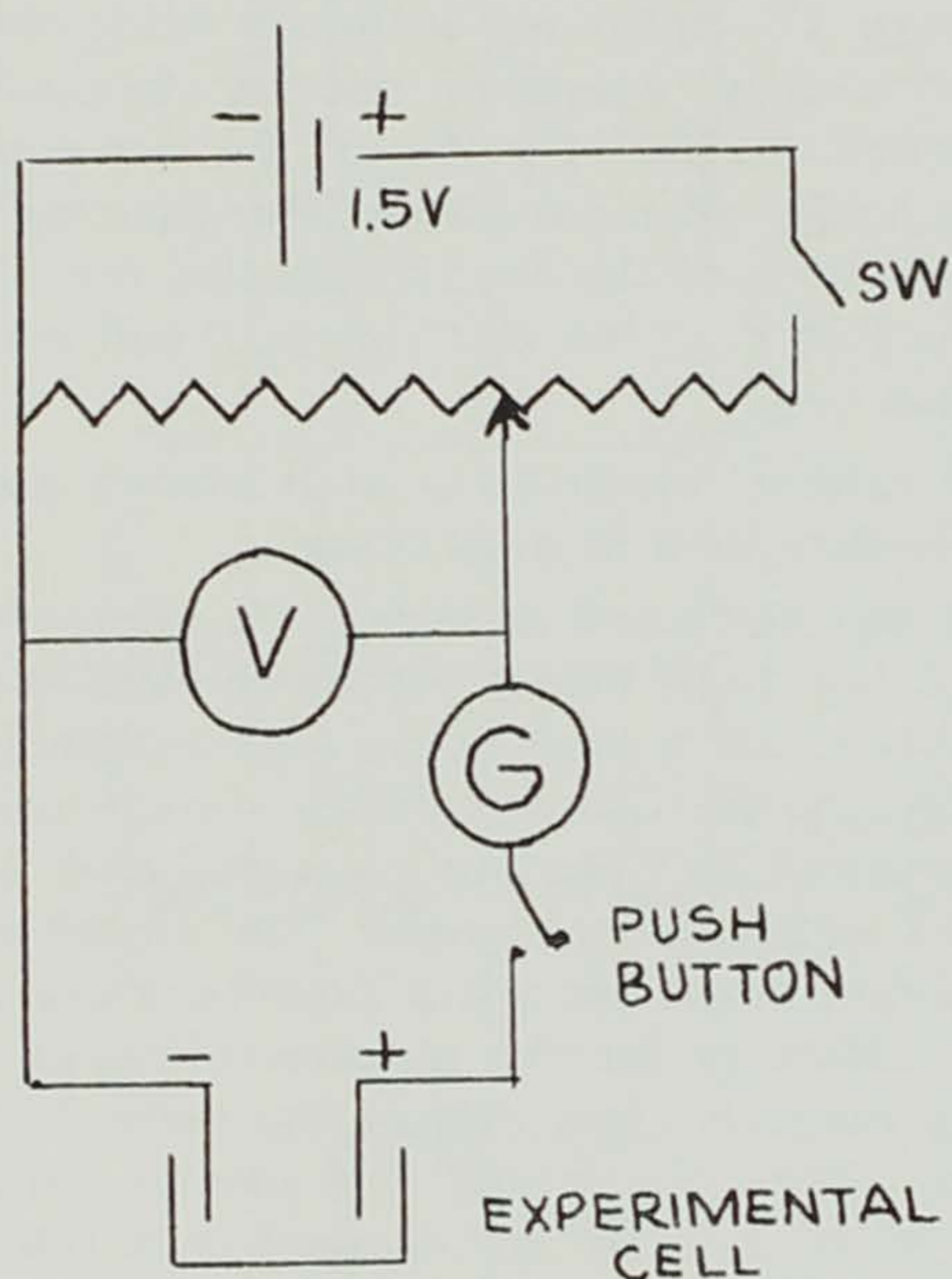


Figure 24

insulated wire to one end of the antimony stick and seal the stick into a piece of 6 mm pyrex tubing with sealing wax. The finished electrode should appear as in the diagram, FIGURE 25.

Making the Calomel Electrode

Directions may be found for this in most laboratory manuals. If a regular glass "half-cell" vessel is available, this should be used. If not, a fairly satisfactory vessel can be made (following the directions below) from three pieces of glass tubing

arranged as shown in the diagram. The main thing is to have the lead wire of platinum sealed in glass in contact with metallic mercury, which is in contact with a mixture of mercury and calomel, which is in turn in contact with a saturated solution of potassium chloride. This KCl solution is then placed in contact through a siphon, a wooden plug, or a ground glass joint with the solution to be measured. If the above conditions of the chemicals are met, the form of the electrode vessel is immaterial.

Making the calomel electrode parts. The platinum wire is first sealed into the end of a piece of 4 to 5 mm O.D. glass tubing. To do this you must obtain a hot pointed gas flame, either from a glass blower's torch or from a portable propane gas torch commonly available from local hardware stores. A Bunsen burner flame will not do the work. Heat a 25 cm length of the small tubing about 10 cm from one end, rotating it in the flame as it is being heated to give it a uniform temperature. When the tube is red hot and soft, remove it from the flame and immediately draw the heated portion to form a fine capillary. When this is cooled, break off the capillary close to the place where the longer section of tubing begins to shrink down. This should leave a piece as shown in FIGURE 27 a total length of about 15 cm.

Hold the short piece of platinum wire in a pliers and insert it into the capillary hole. Hold both tube and platinum wire in the flame just beyond the point until the glass fuses around the platinum making a tight seal as in FIGURE 28.

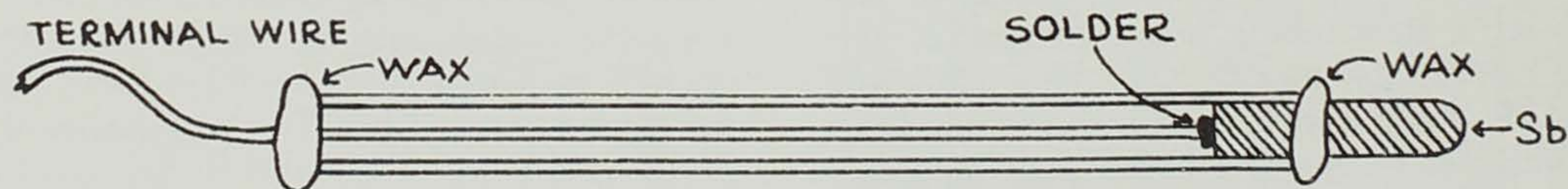


Figure 25

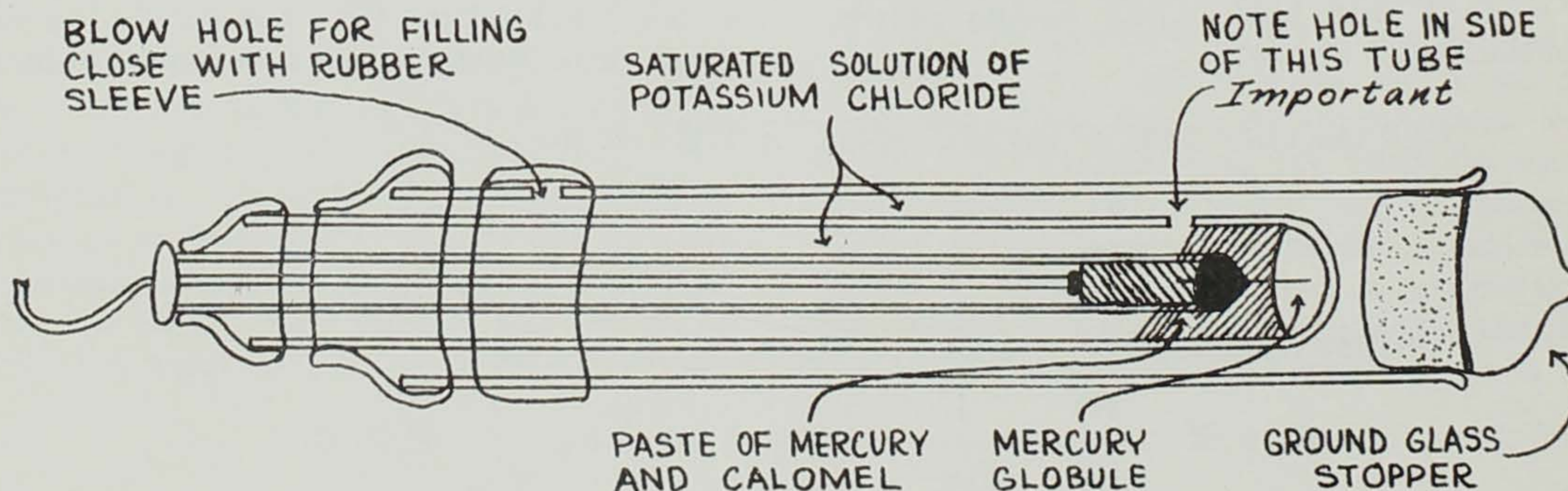


Figure 26

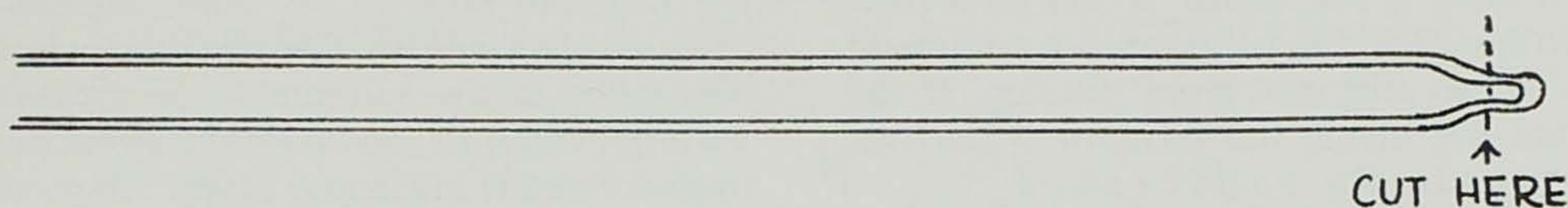


Figure 27

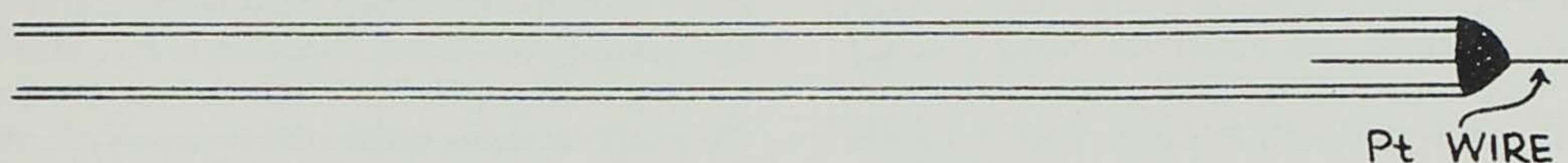


Figure 28

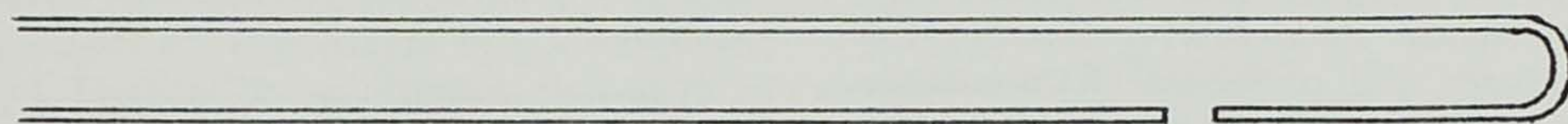


Figure 29

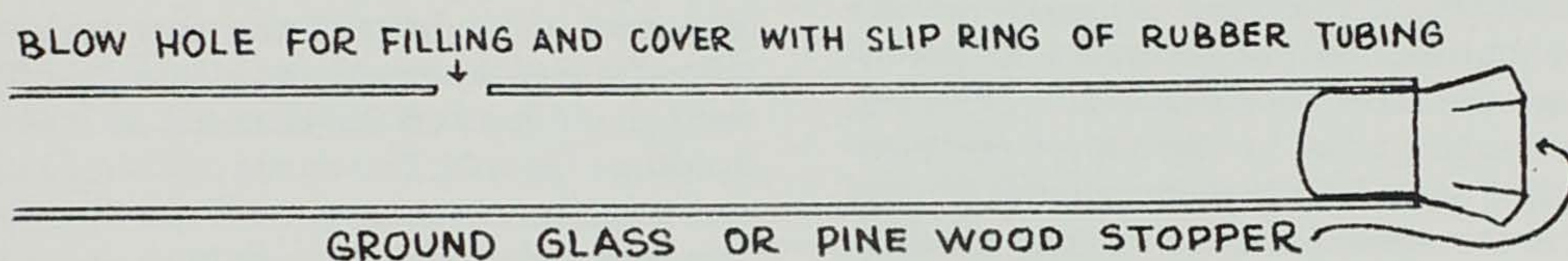


Figure 30

While this is cooling, put a piece of Wood's metal or other low-melting alloy in a vise and saw it with a clean hacksaw, catching the sawdust in a piece of paper. Pour the metal dust into the electrode just made and jar it down so that it settles in a layer $\frac{1}{2}$ inch deep around the platinum wire. Skin the insulation from an 18-inch length of stranded electrical wire and scrape this clean and bright, and run it down the tube into the Wood's metal powder. Place this assembly in a beaker of water and bring the water to a temperature which will melt the Wood's metal. Remove as soon as the metal has melted down and allow it to cool. This gives a good electrical connection which is permanent. If no Wood's metal is available, a globule of mercury can be used; but it is messy and likely to be spilled, and the permanent connection is better.

Cut a 15 cm length of slightly larger glass tube, 7 to 8 mm O.D., into which the electrode tube will slip with a very loose fit. Seal the end of this tube in the flame by pulling out a capillary, melting this off, and then blowing slightly to round off the end. Next blow a very small hole in the side of this tube about 20 mm from the end seal. Do this by heating and blowing in two steps in order to localize the heating and secure a smaller hole. After the glass has blown out, cut off the debris and fire polish the hole so it is flush with the side of the tube as in FIGURE 29.

Cut this tube to a length about 2 cm shorter

than the electrode tube and fire polish the upper end.

Finally, cut a slightly shorter length of 11 to 12 cm O.D. tube to about the same length as the one above. Find a discarded glass stopper about the size of this tube. Heat one end of the tube in a flame for about 15 mm distance at one end, and flare the tube slightly with a carbon rod so that it has the same taper as the glass stopper. Grind the stopper into the tube using 100 mesh carborundum powder or emery and 50 per cent glycerine-water solution as a lubricant. A reasonably good but not exact fit is required. See FIGURE 30. If the stoppered tube will hold water without dripping, it is satisfactory. Cut off the handle of the stopper. Fire polish the other end and assemble the entire calomel electrode as described and shown in FIGURE 26.

NOTE: If a glass stopper cannot be ground in, substitute a close fitting plug of soft pine. A cork or rubber stopper is *not* satisfactory as K^+ and Cl^- ions must be able to migrate through the plug or around it.

Assembly of the electrode. Prepare a saturated solution of potassium chloride by adding 50 g of KCl C.P. 100 ml of H_2O . Shake and allow to come to room temperature. Allow the undissolved solid to remain in contact with the saturated solution. Place a small globule of pure mercury in a small mortar, add about 5 g of mercurous chloride Hg_2Cl_2

and moisten with a few drops of the potassium chloride solution. Grind together until a gray paste is obtained. If too stiff, add more solution. If the mercury does not blend, the mixture is too wet and a little more Hg_2Cl_2 should be added.

To the second glass tube made, that with the sealed end, add a small globule of mercury. Transfer the gray paste to this electrode using a small spatula or wooden paddle. Finally, fill to the side opening with saturated KCl solution. Then take the electrode tube with the platinum contact and push the contact end down inside the tube above until the platinum makes contact with the mercury globule. Disregard the paste and KCl solution forced out through the hole in the side; but once the interior electrode is in place, avoid raising it as this introduces air bubbles. Finally, fill the outer tube about half full of saturated KCl solution and place the inner assembly inside this. The spaces between the tubes at the top of the electrode should be sealed with short lengths of rubber tubing if the electrode is to be kept for any length of time. The completed electrode should appear as in FIGURE 26.

Potentiometer

Assemble the electrical components as shown by the circuit diagram. Connections should be soldered or made with solderless couplings giving dependable connections. If a 0-1 voltmeter is used, the circuit is exactly as shown. If a milliammeter is used as a voltmeter, it must have a resistance in series with it as a "multiplier." The number of ohms of the resistor will depend on the milliammeter used but can be figured from Ohm's law. For example, if the milliammeter reads from 0-1 ma for full scale, and you want it to read 0-1 volt, the total resistance of milliammeter and multiplier must be such that when 1 volt is placed across the combination, exactly 1 ma current will flow:

$$I = 0.001 \text{ amp} = \frac{1 \text{ volt}}{1000 \text{ ohms}}$$

$$R_{\text{meter}} + R_{\text{multiplier}} = 1000 \text{ ohms}$$

When the potentiometer is assembled, test as follows: Turn the potentiometer dial to counterclockwise position. Turn battery switch to "on" position. Both voltmeter and galvanometer should read zero. Move dial in a clockwise direction, and the needle on the voltmeter should rise reaching 1 volt when the dial has been turned about two-thirds of its total rotation. The galvanometer will remain at zero if no electrodes are connected or if the push button is not depressed.

Operation. Connect the antimony electrode to

the negative pole of the potentiometer and the calomel electrode to the positive pole. Dip both electrodes in the solution to be measured. Turn battery switch to "on" position. Press push button until contact is made and rotate the potentiometer resistor until the needle of the galvanometer returns to zero. Read the voltmeter. Enter the voltage reading on the calibration graph and read the corresponding pH from the curve. To measure a different solution, rinse off electrodes* and replace in the next solution, repeating procedure above. It is a good plan to stir the solution while making a reading.

Calibration. This electrode system is theoretically accurate, and the pH should be calculable from the Nernst equation shown on page 71. However, the purity of the chemicals, especially that of the antimony in the electrode, often throws the calculations off by as much as a several hundredths of a volt. The curve of voltage as ordinate versus pH as abscissa, however, will still be linear. If a calibration curve is established by measuring the voltage produced by buffers of known pH, the system will be accurate for practical purposes. This is essentially the procedure used with glass electrode pH meters. To calibrate, only two buffers are needed to establish the straight line curve. One should be around 3 or 4 pH while the other should be above pH 7. If buffer solutions are not already available, the following are suggested.

Ten g glacial acetic acid + 10 g sodium acetate in approximately 200 ml H_2O gives a pH of 4.73 at 25° . A saturated solution of acid potassium tartrate gives a pH of 3.55 at 25° . A 0.05 molar solution of acid potassium phthalate gives a pH of 4.00 at 25° . A 0.05 molar solution of sodium bicarbonate (baking soda) gives a pH of 8.2 at 25° . Other buffer formulas may be found in chemistry handbooks. If after long use or standing idle, the antimony electrode goes "off calibration" as determined by checking against the known buffers, one needs only to scrape the metal clean with a knife blade to restore its activity.

EFFECT OF SOLUTES ON THE BOILING POINT OF A SOLUTION—PUPIL EXPERIMENT

Materials. 6mm glass tubing
thermometer, 0.5° or 0.1° divisions
with a range of -10° to $+110^\circ \text{ C}$.

* If a wood plug is used in place of the glass stopper, the plug may be contaminated with very acid solutions or very alkaline solutions and should be replaced frequently.

test tube, 25 × 150 mm

urea

dextrose

glycerol

Discussion. The boiling point of a liquid is defined as the temperature at which the vapor pressure of the liquid is equal to the atmospheric pressure. Dissolving a nonvolatile solute in a solvent lowers the vapor pressure of the solvent. The boiling point of the resulting solution is higher than that of the pure solvent.

The purpose of this experiment is to determine the effect of nonvolatile solutes on the boiling points of water and other solvents.

The amount of elevation of the boiling point depends upon the concentration of the solute. For ideal dilute solutions, the change in boiling point is proportional to the concentration of the solute. Mathematically, this can be expressed as $T_b = K_b m$. T_b represents the elevation of the boiling point of the solution above that of the pure solvent. The constant K_b , the molal boiling point constant, is characteristic of each solvent. The quantity (m) represents the molality of the solution or moles of solute that are dissolved in 1000 grams of the solvent. One may determine the molal boiling point constant and the molecular weight of a nonvolatile solute if the change in boiling point is determined for a weighed quantity of solute dissolved in a known amount of solvent.

When determining the boiling point of a solution, the thermometer must be in contact with the solution. Superheating, heating a solution above its boiling point, may occur. In such cases, the observed temperature is not the true boiling point of the solution. Superheating is caused when bubbles of a vapor are not readily formed in the solution.

To avoid superheating, a special vapor pump, Cottrell pump, is used. While elaborate Cottrell pumps are available, a suitable one can be constructed from a piece of glass tubing or a semimicro funnel tube.

Procedure. (a) Construction of a simple Cottrell pump.¹

Seal the end of a 20 cm piece of 6 mm glass tubing. Blow a thin bulb on the sealed end. Break or cut the bulb from the tubing and fire polish the flared end. Approximately 7 cm from the flared area make a 90° bend. Cut the glass tubing in two

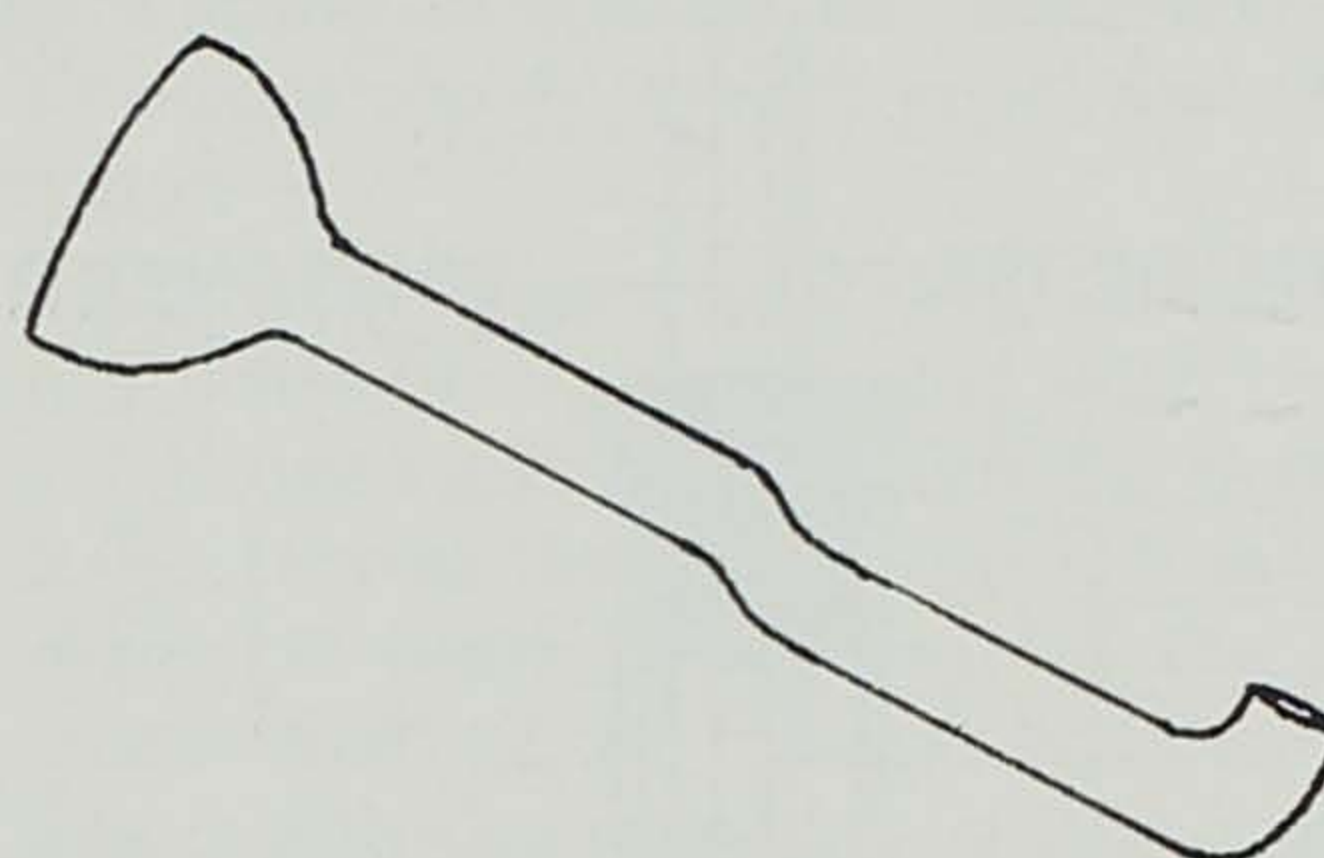


Figure 31

in the middle of the bend. The pump should resemble the drawing.

A semimicro funnel tube is also suitable for construction of a simple Cottrell pump.

(b) Setting up the apparatus. Wind one-half of a 5 cm length of copper wire around the stem of the pump, extend the wire remaining above the outlet of the pump and wind the rest of the wire around the thermometer so as to hold the outlet of the pump against the thermometer bulb. Place the thermometer and pump assembly, flared end down, in a test tube. On boiling the liquid, the pump should reflux the liquid against the bulb of the thermometer. See FIGURE 32.

(c) Determining the boiling point of water by use of the apparatus in (b).

(d) Weigh the test tube to the nearest 0.01 of a gram. Add about 10 to 15 ml of water and weigh again. Weigh separately about 3.00 to 4.00 grams of urea or dextrose and dissolve in the water. Determine the boiling point of the solution. The temperature should be recorded shortly after the liquid starts refluxing.

NOTE: The instructor may wish to prepare several different concentrations of solutions for class use. Electrolytes such as sodium chloride may be used.

(e) It is possible to use other solvents besides water. Carbon tetrachloride is often used because of its rather high molal boiling constant of 5.03° at 760 mm. A correction of 0.0013° per mm difference from 760 mm may be made. Suitable solutes for carbon tetrachloride are naphthalene, camphor, or benzoic acid.* About 15 ml of carbon tetrachloride with three grams of solute should be used.²

(f) Extreme care should be taken when using carbon tetrachloride because of its toxic nature. This experiment should be carried out in the hood. It should not be carried out as a group experiment

¹ C. S. Hoyt, "Boiling Point Elevation as an Experiment in General Chemistry," *Journal of Chemical Education*, XXIV (1947), 563.

* NOTE: Have the pupil account for abnormal results with benzoic acid.

² Hoyt, *loc. cit.*

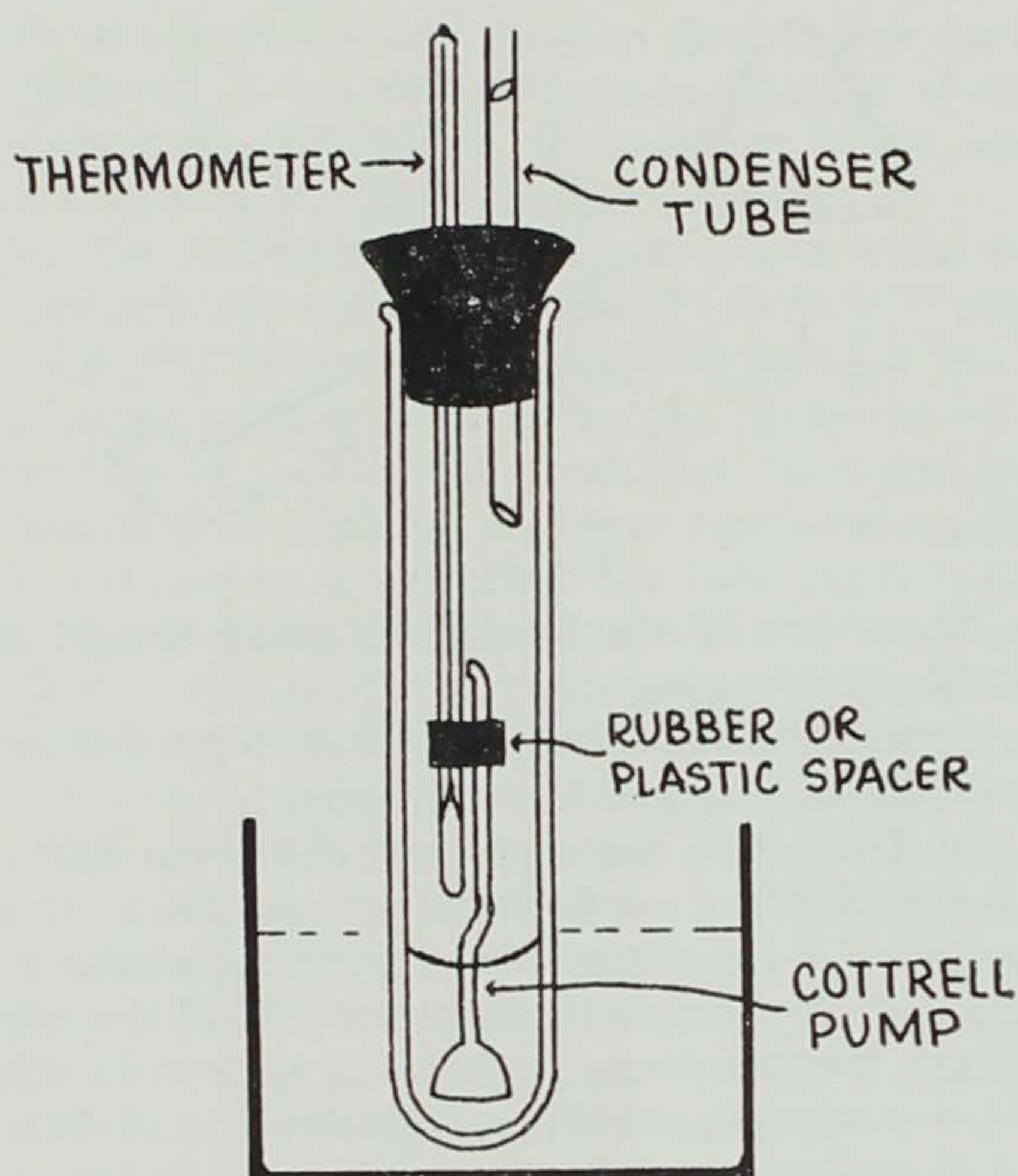


Figure 32

in the open classroom. The recommended apparatus is included below. Do not heat the water bath excessively, just enough to boil the carbon tetrachloride solution. The carbon tetrachloride vapors should condense in the condenser tube just above the rubber stopper. Take several temperature readings, and when the temperature is constant, record the reading to the nearest 0.1 of a degree.

Calculations. (1) Using the weight of urea, the molecular weight of urea, and the weight of water, compute the molality of the solution. (2) From the molality of the urea solution and the observed elevation of the boiling point, determine the molal boiling point constant of water. How does your value compare with the theoretical molal constant? (3) Use the theoretical molal boiling constant of water to determine the molecular weight of the solute, either of urea or of an unknown.

Questions

1. Account for the elevation of the boiling point of water when a nonvolatile solute is dissolved in water.
2. If the boiling, refluxing of the solution against the bulb of the thermometer, is continued for several minutes, what effect would this have on the boiling point of the solution? How could this source of error be reduced?
3. Why must the vapor pressure of the solvent or the solution be slightly above the atmospheric pressure if boiling is to begin?
4. Pupils doing these experiments on different days may obtain slightly different results. Account for this variation.

CONSTRUCTION OF AN OZONIZER— PUPIL PROJECT

An ozonizer is an apparatus for the production of ozone from oxygen, either from the air or from the pure oxygen obtained from a cylinder. It consists of a double-walled glass vessel through which the oxygen is passed and subjected to a high voltage electrical charge. See FIGURE 33.

- Materials.**
- 1 1 1/4 in. × 8 in. side arm test tube
 - 1 1 in. × 6 in. plain test tube
 - 1 rubber stopper to fit smaller tube
 - 1 short length of 6 mm glass tube
 - tin foil
 - shellac or diluted airplane cement
 - model airplane cement
 - spark coil (induction coil as from Model T Ford or as obtained from supply houses for demonstrations, etc.)
 - wires for connections

Directions. Cut a three-inch length of the 6 mm glass tubing and fire polish both ends. Place in rubber stopper and connect outer end to a length of rubber tubing. Attach this to the mouth of the

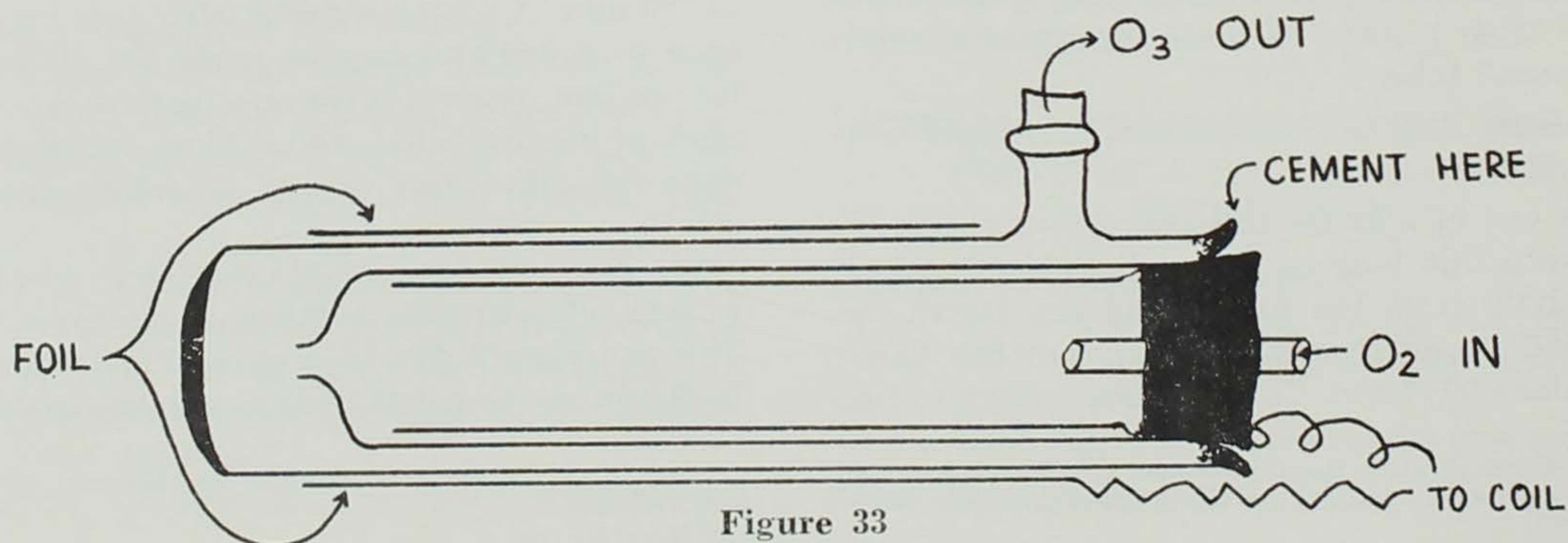


Figure 33

smaller test tube. Heat the closed end of the test tube in the pointed flame of a glass working torch or a small flame from a bottled gas (propane) soldering torch. When the glass begins to glow, remove from the flame and blow into the rubber tubing until a small blister extends from the test tube. Return to the flame, reheat the blister and blow out the glass entirely, leaving a small hole a few millimeters in diameter at the end of the test tube. Fire polish this.

Coat the inside of this tube with a layer of tin or lead foil, allowing a strip of foil to extend from the mouth of the tube for an electrical connection. It is better to cement the tin foil into place with diluted model airplane cement. Place this tube inside the larger tube bearing the side arm. Hold the assembly vertical so the tubes will center, and cement the lips together with model airplane cement so a gas tight seal is obtained. Finally, coat the outside of the outer tube with a layer of tin or lead foil, cementing it into place. Attach thin wires to both tin foil coatings, and cement the wires in place so they will not tear the tin foil. Replace the stopper in the inner tube. The entire apparatus should appear as in the diagram.

Operation. The tube leading to the inner test tube is the inlet and should be connected to the source of oxygen. The side arm of the large tube is the outlet for the mixture of ozone and oxygen. Connect the two wires to the spark coil and turn on the current. Pass oxygen or air through the apparatus, and immediately, the penetrating characteristic odor of ozone will be observed. OZONE IS POISONOUS—USE GOOD VENTILATION—DO NOT BREATHE IN ANY BUT MINUTE QUANTITIES.

CATALYTIC OXIDATION—PUPIL PROJECT

This is a project suitable for class use in connection with the study of oxygen, of combustion, or catalysis as related to the velocity of chemical reactions.

Secure a "Jon-ee" or other hand warmer depending upon the catalytic oxidation of a hydrocarbon for its heat.* Follow the directions for fueling the warmer and getting it started. Vary the amount of air admitted to the combustion chamber and relate this to the amount of heat produced. Bring the warmer into a dark room and observe the glow of combustion on the catalyst surface.

* Very inexpensive warmers are obtainable for class use at one dollar each from Edmund Scientific Company, Barrington, New Jersey.

If the warmer is considered "expendable," the group of pupils should take it apart and study its construction.

Reassemble the warmer, light it, and then place it in an atmosphere contaminated with hydrogen sulfide. Hydrogen sulfide is a poison for many catalysts. Is this true in the present case? That is, does the combustion cease as a result of the catalyst being inactivated? If the catalyst has become poisoned, remove the material, place it in a crucible, and heat to a high red heat with a Bunsen burner or better, a blast lamp. Replace in the warmer and see if it will now light. What has taken place to reactivate the catalyst?

The catalyst in warmers is usually platinized asbestos. This can be made by soaking asbestos fiber in a solution of chloroplatinic acid, H_2PtCl_6 , and igniting in a crucible as was done above in reactivating the poisoned catalyst. H_2PtCl_6 can be made by dissolving scraps of platinum in aqua regia.

See if the platinized asbestos catalyst will relight the gas issuing from a well-adjusted Bunsen burner. Repeat with an alcohol lamp. If the catalyst works well, secure some small aluminum tubes such as those in which Kodachrome film is marketed, and see whether you understand the principle of the hand warmer well enough to construct your own.

THE DETERMINATION OF CARBON DIOXIDE IN A CARBONATE—PUPIL PROJECT†

This is a simple project involving the procedures of quantitative analysis. All apparatus, with the exception of the balance used in weighing, can be easily made from common laboratory glassware.

- Materials.**
- 2 drying tubes 150 mm long \times 15 mm diameter
 - 1 drying tube 100 mm long \times 15 mm diameter
 - 1 reaction flask, 125 ml Erlenmeyer
 - 1 acid reservoir vial 64 \times 15 mm
 - 1 baffle tube 65 mm long \times 12 mm diameter (test tube)
 - glass tubing 3 mm O.D. and 5 or 6 mm O.D.
 - 1 rubber stopper #00 two hole
 - 2 rubber stoppers #0 one hole
 - 1 rubber stopper #4 one hole

Assembly. Assemble the components as shown in FIGURE 34, and test for airtightness. The rub-

† Used by permission of Lester Dalton and J. L. Moriarty, based on a paper read at Iowa Academy of Science.

ber stoppers should be of fresh lively rubber and should be boiled in 10 per cent sodium hydroxide solution until all "bloom" has been removed. Drying tubes A, B, and the baffle tube in C should be filled with anhydrous calcium chloride, granular, free of dust, and of a size to pass through a 10 or 12 mesh per inch sieve. It is best to pass a small amount of carbon dioxide through tube C prior to use of the apparatus in a determination to react with the traces of calcium hydroxide sometimes contaminating calcium chloride.

Directions. Weigh flask E dry and empty. Add about 5 grams of dried baking powder, or 1 gram of calcium carbonate or marble, or other sample to be analyzed, to the flask and weigh again. Add just sufficient freshly boiled distilled water to the sample to seal the lower end of the tube F when it is in place. Add sufficient 1:1 hydrochloric acid solution to nearly fill the acid reservoir G. Attach the sample flask to the apparatus but leave drying tubes A and B and connecting tube H off for the time being. Weigh the entire apparatus. A similar "dummy" apparatus may be used as a tare. Attach drying tubes A and B and gently draw enough air through the system to transfer the acid from its reservoir to the sample flask. As soon as the sam-

ple has decomposed and bubbles of carbon dioxide are no longer being liberated, heat the solution in the flask until warm, but do not boil. At the same time, draw a gentle stream of air through the entire system by means of an aspirator. It is best to have a bubble counter between the drying tube B and the aspirator and to restrict the flow of air to two bubbles per second or less. Maintain the flow of air until the apparatus cools to room temperature, or about 15 minutes. Weigh the apparatus as before with drying tubes A and B and connecting tube H detached.

The loss in weight of the apparatus (difference of the two last weighings) divided by the weight of the sample (difference of the first two weighings) multiplied by 100 gives the per cent of carbon dioxide in the sample. Make a second determination to check your results.

Questions

1. What is the purpose of the baffle tube in C?
2. Why is drying tube A necessary? drying tube B?
3. Why is it desirable to use a much larger sample of baking powder than of limestone?
4. Explain what is meant by relative error. How does this relate to question 3?

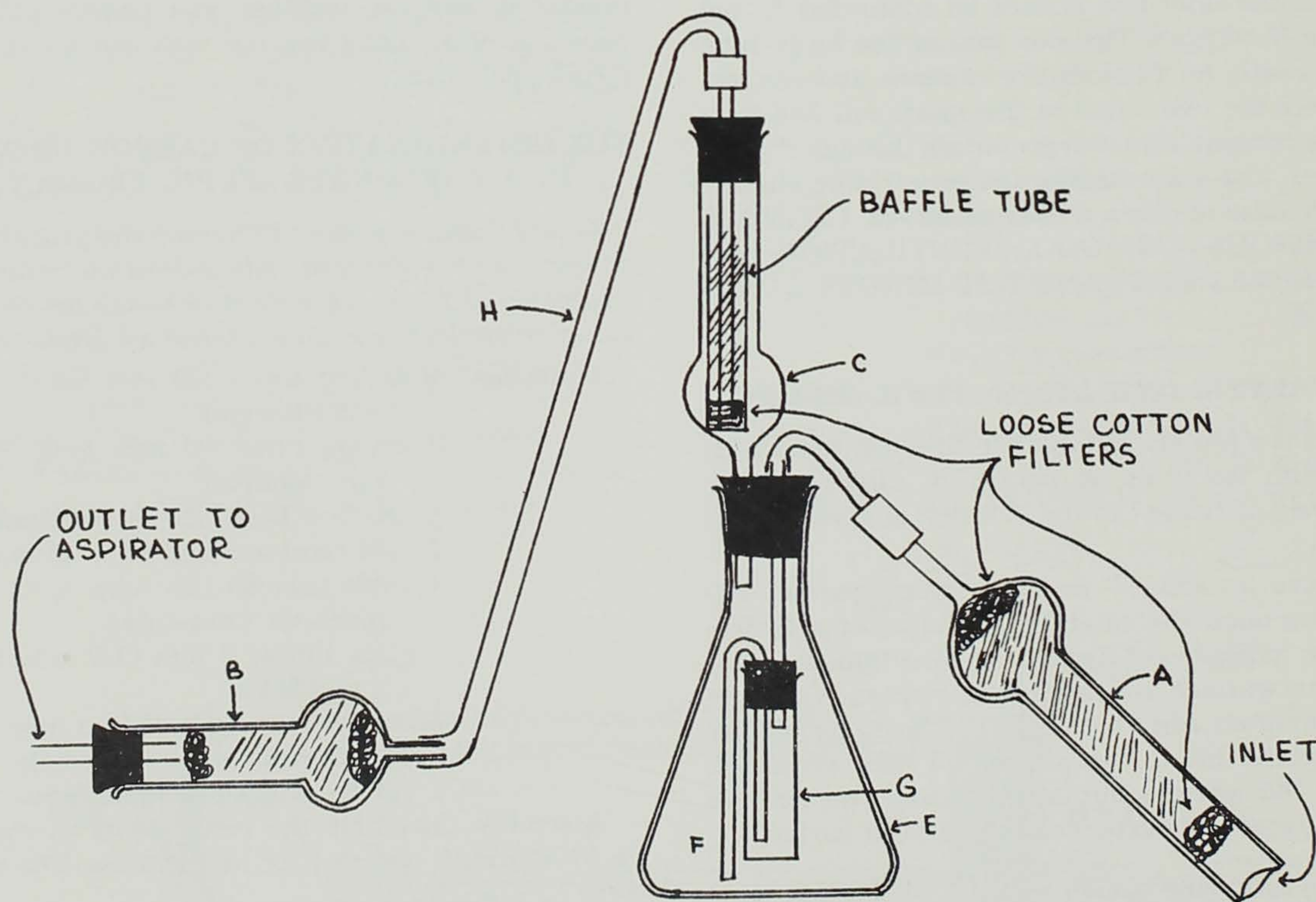


Figure 34

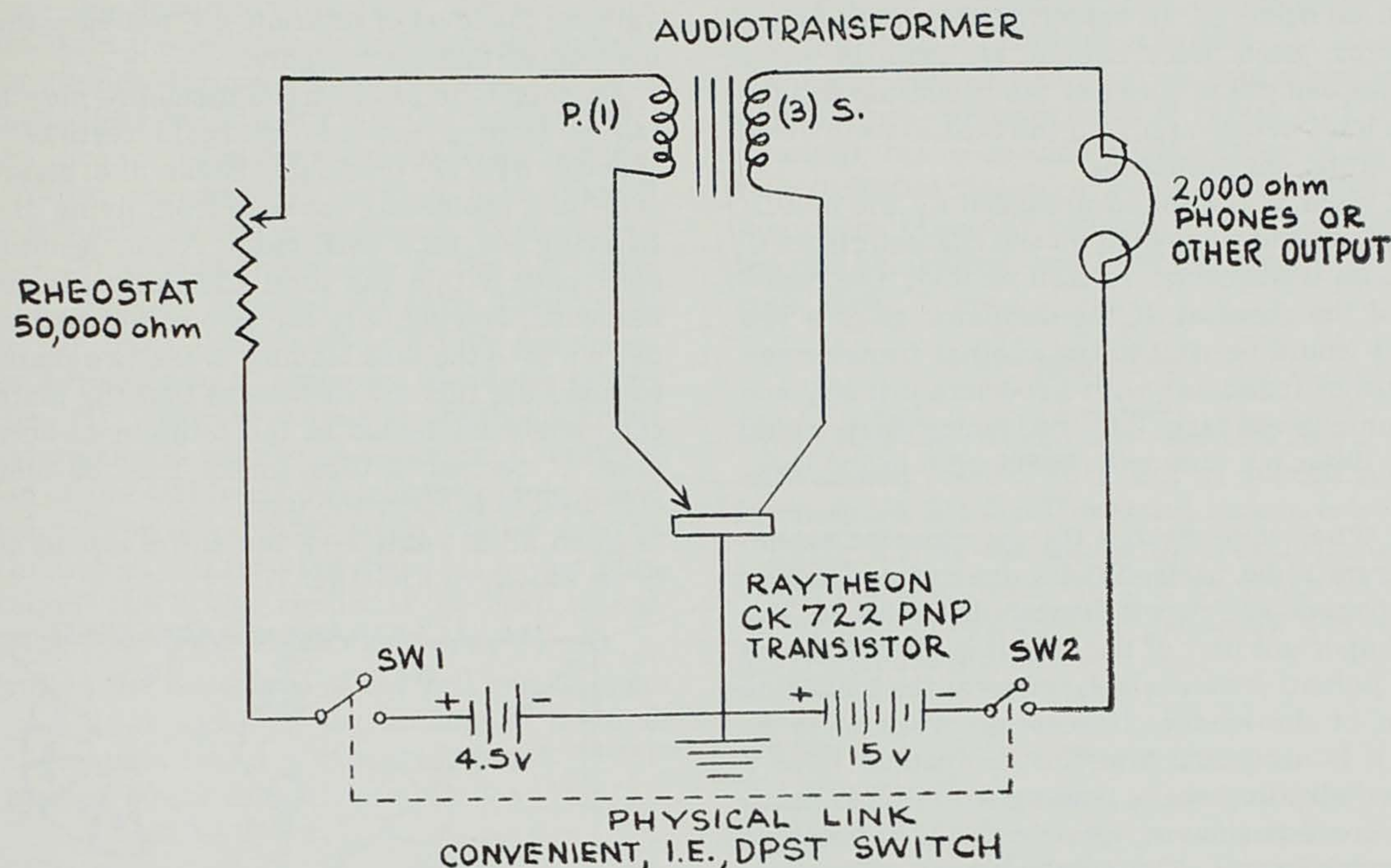


Figure 35

5. Why is it advisable to warm the solution in freeing it from dissolved carbon dioxide? State the general principle involved.

6. Why is it necessary to dilute the acid before using it in the apparatus?

7. Why should the calcium chloride be free of dust?

8. Why are there plugs of porous cotton placed in the drying tubes before filling with calcium chloride?

9. If the calcium chloride in tube A was wet and not drying the air as effectively as that in tube C, would your determination have an error that was above or below the true value?

Suggestions for further work. The method used in this determination is that of evolution of the CO_2 and determination of the weight loss. Carbon dioxide can also be determined by absorption and by precipitation. Devise apparatus for each of these methods and write out directions for their use. Test one of these methods to see whether it is more or less accurate than the one described above.

Could the apparatus described be simplified further? Could it be improved by increasing its complexity? If you have ideas for either, write them up and try them.

ELECTROLYTIC CONDUCTIVITY APPARATUS—PUPIL PROJECT

A project for the development of a low cost conductivity apparatus suitable for conductometric titrations or for the determination of the conductivities of electrolytes with moderate accuracy.

Electrolytic conductivity measurements require: (a) a source of audio frequency alternating current of about 1000 cycles per second, (b) a Wheatstone bridge for actual measurement, (c) a detection device to indicate the point of balance, and (d) a cell with electrodes in which the solution to be measured is placed.

For (a) a simple transistor oscillator is satisfactory, although an electron tube oscillator may be made. Louis E. Garner gives a suitable transistor oscillator operating from low power batteries.¹ This circuit was originally published in *Radio and Television News*. The circuit is as in FIGURE 35.

Few specific directions for the assembly of the oscillator circuit are needed for those familiar with radio construction. The various components with values as marked on the circuit diagram should be assembled as indicated. All connections

¹ Louis E. Garner, *Transistors and Their Application* (Chicago: Educational Book Publishing Division, Coyne Electrical School, 1954), 86.

should be soldered. In soldering the wires of the transistor, each wire should be gripped by a needle-nosed pliers between the transistor and the solder joint before applying the heat so as to avoid overheating the transistor.

The pitch of the sound produced by the oscillator will depend somewhat on the characteristics of the audio transformer as well as upon the resistance of the rheostat. If the oscillator gives a low note, it would be well to try another transformer. The author found one with a 500 ohm primary and 6000 ohm secondary D.C. resistance that would give a 1000 c.p.s. note with 50,000 ohm in the rheostat.

The Wheatstone bridge (b) may also be assembled from radio parts. Trial values of resistances are shown on the circuit diagram FIGURE 36. The most important part of the circuit is the 3000 ohms linear potentiometer which serves as the balancing system of the bridge. If accurate work is to be done, it is necessary that this component have a dial or indicating scale, reading in uniform units, which corresponds very closely to the rotation of the potentiometer. The total number of graduations on the scale is unimportant, but if the potentiometer rotates 270° from minimum to maximum setting, the scale on the dial must also cover this same number of degrees of rotation. In attaching the dial to the shaft, it is best to set the moving contact of the potentiometer to the exact midpoint of resistance and have this indicated by the middle mark of the dial scale. If the ends of the scale are a little "off calibration," this is of less importance than that the midpoint be exact.

The detection device (c) used to determine the point of balance of the bridge may be a pair of radio headphones, preferably of good quality and of about 2000 ohms resistance. Since the point of balance corresponds to a *minimum* of sound in the headphones, a simple one transistor or one tube audio amplifier is advantageous. Circuits of such amplifiers may be found in radio publications or may be obtained by writing to manufacturers of transistors.

The cell (d) used for the actual measurement may be made from two stainless steel plates approximately 1 cm square and spaced from 0.5 cm to 1 cm apart by means of polystyrene or similar nonporous plastic. They should be so designed that they can be suspended in a beaker of the solution to be measured and so that exactly the same area of surface is in contact with the liquid during each measurement. The wires leading to the electrodes should be of relatively small size, No. 20, so that

varying the level of solution will not vary the conducting surface appreciably.

Another type of electrode assembly may be obtained from a burned out radio rectifier tube: type-80, 5Y3GT, or 5U4G. Break the glass bulb carefully, protecting the eyes from flying glass by covering the tube with cloth. Avoid bending the electrodes within the tube. Carefully remove the filaments, leaving only the two nickel plates. Solder wires to the pins leading to the two plates and suspend the tube by its base so that the plates are completely immersed in the solution to be measured. If desired, a tube socket may be mounted and used to support the tube.

When all is assembled, the entire circuit should be as shown in FIGURE 37.

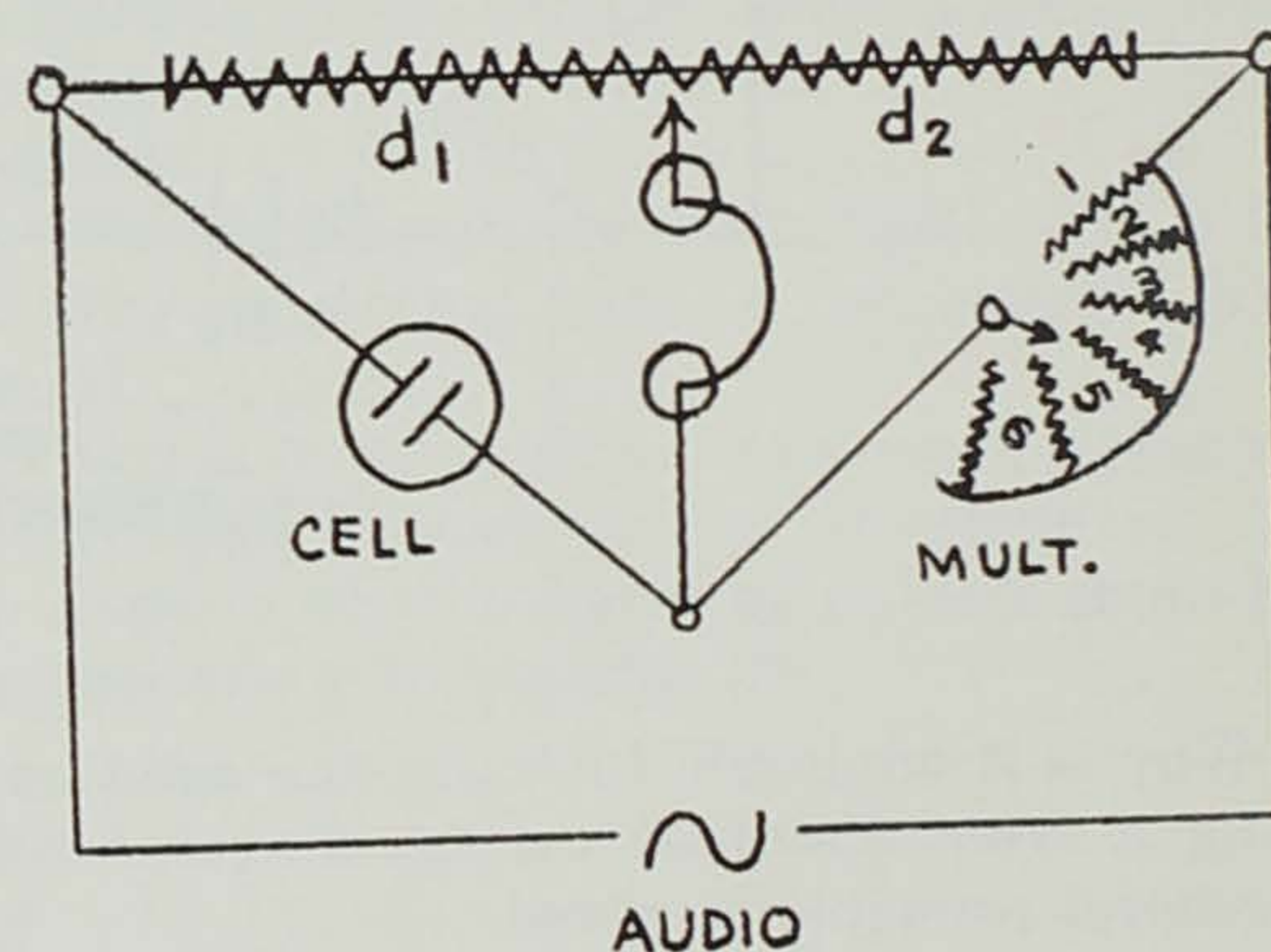


Figure 36

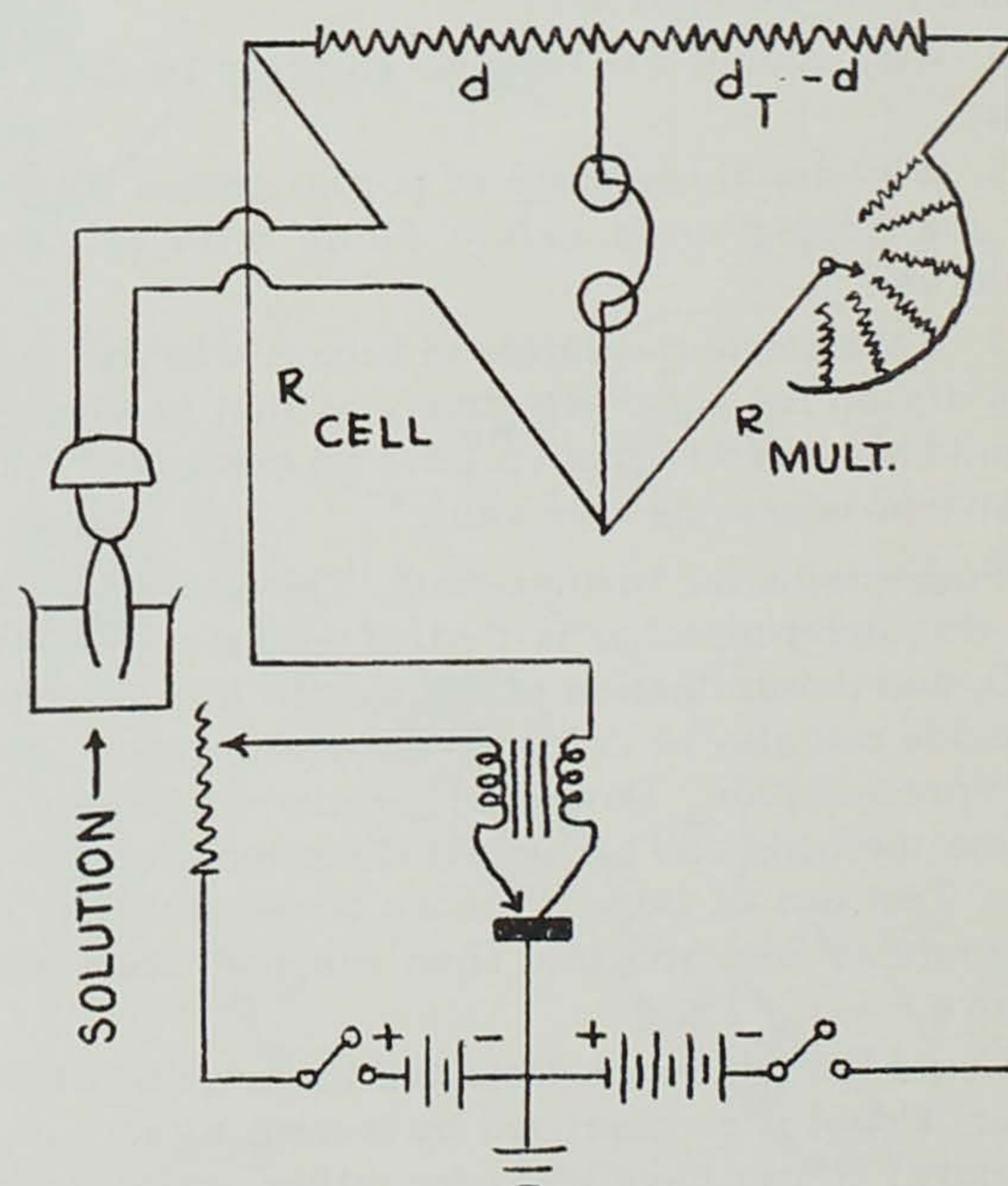


Figure 37

To use the apparatus, immerse the electrodes completely in the solution to be measured and see that no air bubbles are clinging to their surface. Close switches Sw 1 and Sw 2 in the oscillator circuit and turn on the audio amplifier if one is being used. Adjust rheostat (1) until a high pitched whine is heard in the headphones. Turn potentiometer (2) so that the dial is at the half scale point, i.e., midway between the ends. Turn range switch (3) until the sound heard in the headphones decreases to a minimum. Readjust potentiometer (2) until the sound in the headphones fades gradually to zero or to a minimum. Further rotation of the dial beyond this point should cause the sound to become louder again. Adjust for the exact minimum and read the dial. It is a good plan to "ground" the bridge circuit at the fixed end of the headphones connection.

Calculate the resistance of the cell and its contents as follows: let d be the reading of the dial where minimum sound is observed, and d_t be the total reading on the dial. Then—

$$d : (d_t - d) :: R_{\text{cell}} : R_{\text{mult.}}$$

Solve this proportion for R_{cell} using as $R_{\text{mult.}}$ the value of the resistance which is in the circuit at this position of the range switch. For example, if the dial on potentiometer (2) is graduated in 100 parts and the dial reads 35 when a minimum of sound is heard and the range switch is set at 10,000 ohms, the proportion would be:

$$35 : (100 - 35) :: R_{\text{cell}} : 10,000$$

$$R_{\text{cell}} = \frac{10,000 \times 35}{(100 - 35)} = \frac{350,000}{65} = 5,400 \text{ ohms}$$

The figure 5,400 ohms is "rounded off" from 5,385 since only two significant figures can be read from the dial and radio resistors such as those used in the multiplier have ± 5 per cent tolerance. One should not fool himself about the accuracy of measurements. Regardless of the arithmetic involved, the actual measurement depends for its accuracy upon all the data entering into the calculation. The final value can be no more accurate than the least accurate measurement included.

The accuracy of this instrument can be improved greatly by using precision resistors in the multiplier and by employing a dial which can be read to three figures.

The specific conductance of a solution is defined as the reciprocal of the resistance of 1 cubic centimeter of solution measured between electrodes 1 sq. cm in area and 1 cm apart. No cell has these exact dimensions. To obtain absolute values, the "cell constant" of the conductivity cell must be

obtained by measuring the conductivity of a known potassium chloride solution.

For 0.02 Normal KCl (1 gram equivalent weight of potassium chloride, or 74.56 grams, in 1 liter of solution) has a specific conductivity of 25°C of $L_s = 0.002768$ reciprocal ohms per centimeter. If R_{cell} is the observed resistance of the solution in the cell being used, and L_s is the specific conductance, then

$$L_s = \frac{K}{R_{\text{cell}}}$$

where K is the cell constant. If 0.02N KCl solution with known value of L_s is used in the cell and R_{cell} is measured, K the cell constant may be calculated.

Using this K for future measurements of unknown solution, their specific conductivities may be calculated from the above formula. The equivalent conductivity then is:

$$\text{eq.} = \frac{1000 L_s}{\text{Normality}}$$

Conductometric titrations are carried out by placing the solution (e.g., a dilute acid) in the cell and measuring its conductivity as described above. Next, some of the titrating solution (dilute alkali) is added from a burette, its volume observed, and after shaking to mix the solutions the new conductivity is measured. This process is continued by adding approximately equal volumes of the titrating reagent at each step until the end point has been passed and approximately twice the volume of reagent to reach the equivalent point (end point) has been used.

The volume of titrating reagent added at each addition should be about 20 per cent of the total

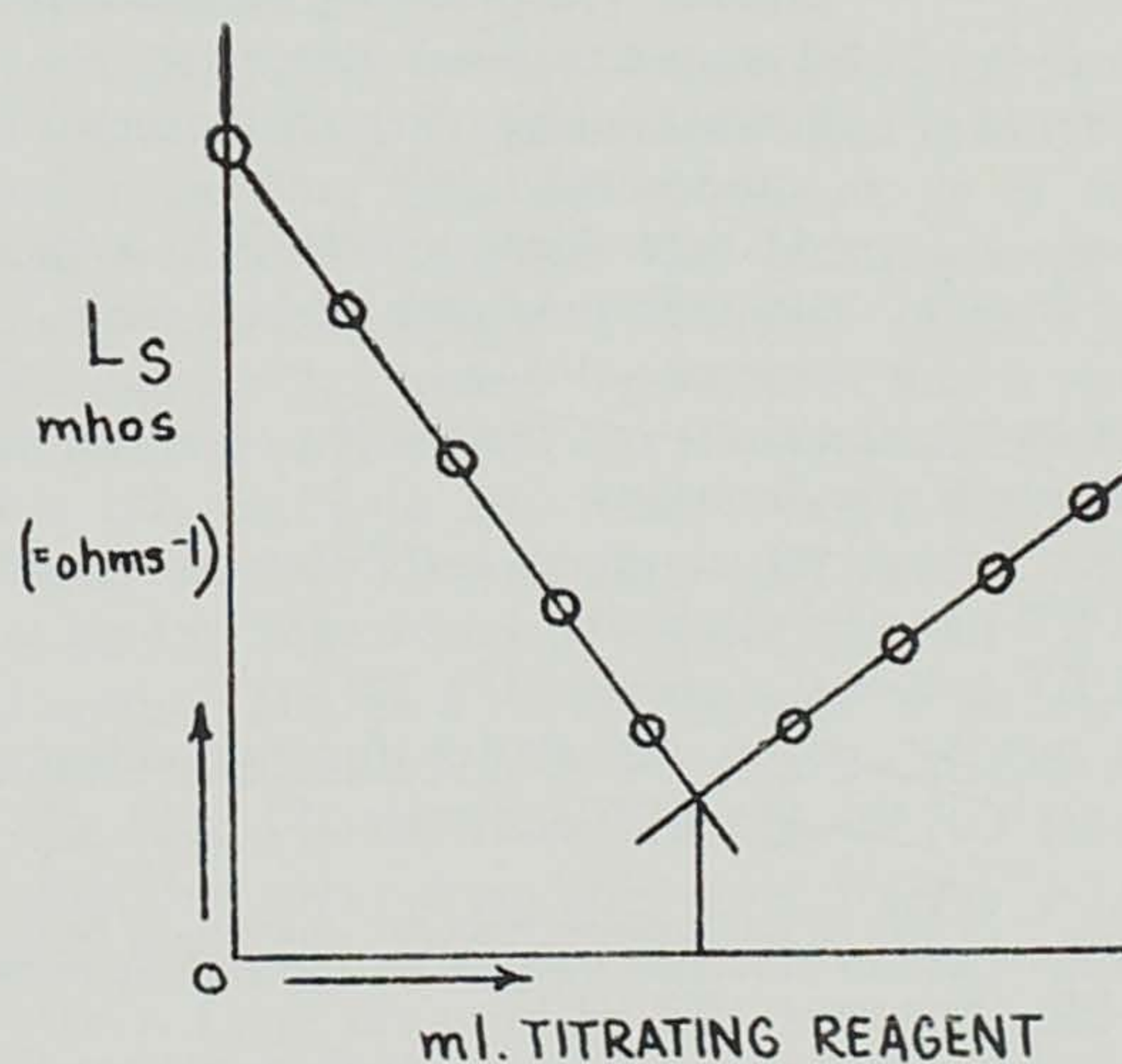


Figure 38

volume necessary to reach the equivalent point. No attempt is made to stop the titration at the equivalent point or to note the exact end point in passing it. Instead, this value is obtained graphically by plotting the various conductivity values against the corresponding volumes of titrating reagent. The end point (equivalent point) of the titration is represented by the intersection of the two straight lines obtained as in FIGURE 38.

Many titrations, not having suitable visual indicators, can be made by this method which is of very wide applicability.

Proctor and Roberts¹ give an additional experiment that can be done using this apparatus.

DEMONSTRATING CONDUCTIVITY OF SOLUTIONS—PUPIL PROJECT

An elaborate device that makes it possible to demonstrate the conductivity of solutions without destroying, transferring, or contaminating them is described in this article.²

Graphite or compressed carbon electrodes wired *in parallel* are placed in 25 × 200 mm test tubes. One lead from each tube is connected to a 12 position selector switch, capable of carrying 1 amp current. By rotating the switch, the conductivity of each solution can be measured by the brightness of a lamp, by a 0-1 ammeter, or by an external projection meter.

Materials. SPST power switch
SPST external meter switch
12 position switch for connecting to solutions recommended J.B.T. SS 14—1 (2-14) position of Mallory 32112J (12 position, non shorting)
0-1 amp a.c.-panel meter
16 watt, clear carbon filament incandescent lamp
grid caps for connecting to carbons
22¼ inch diameter carbon rods
selected solutions

An external circuit for the testing of additional solutions is also included.

Only two of ten standard test tubes are shown.
SPST power switch (1)
SPST external-meter switch (2)
12 position power tap switch for connecting to each of the ten standard cells (3)

¹ James S. Proctor and John E. Roberts, "Analysis of Aspirin: A Conductometric Titration," *Journal of Chemical Education*, XXXVIII (September, 1961), 471.

² Fred B. Eiseman, Jr., "A Device for Demonstrating Conductivity of Solutions," *Journal of Chemical Education*, XXXIII (1956), 445.

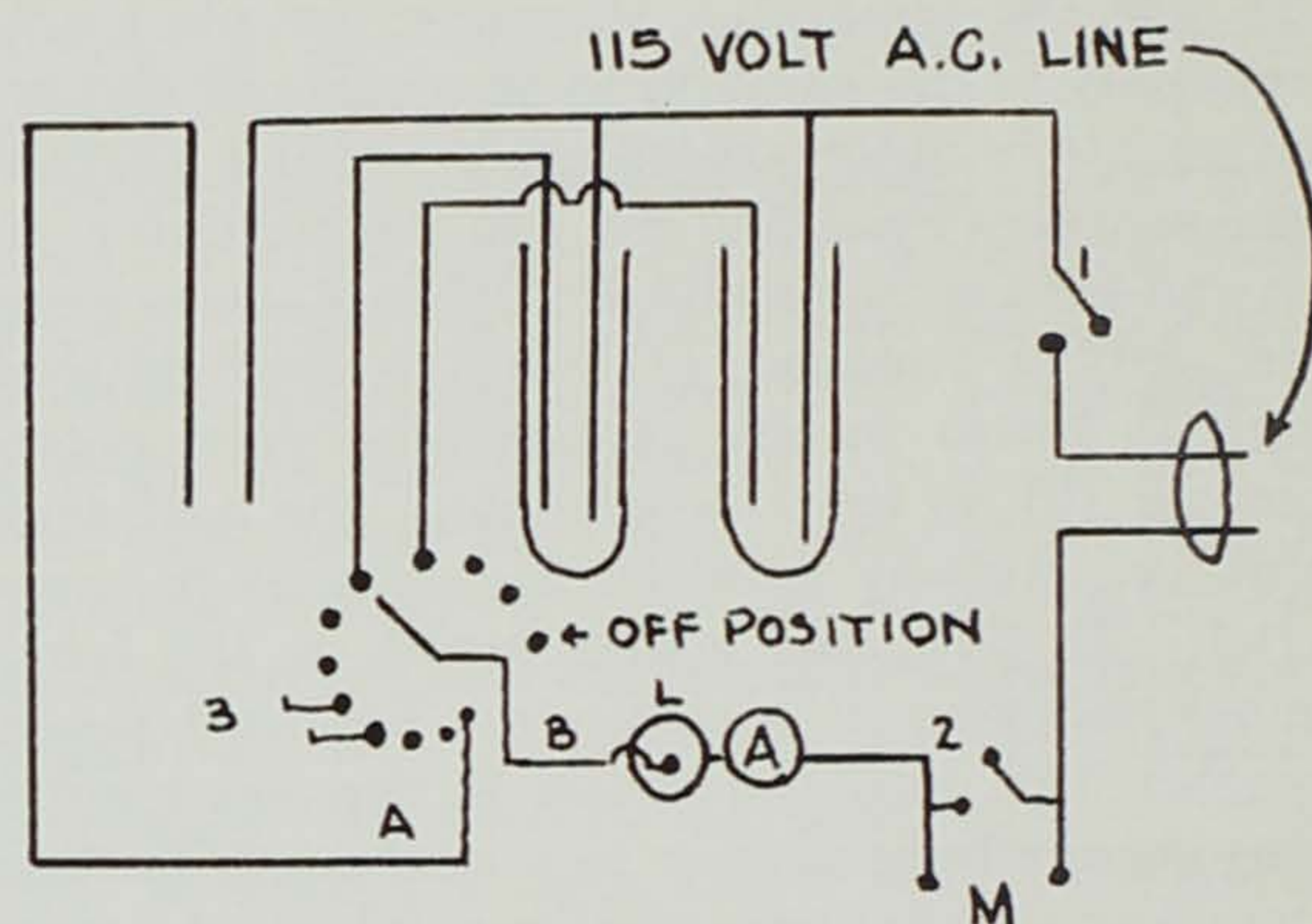


Figure 39

0-1 amp a.c.-panel meter (A)

16 watt, clear envelope carbon-filament incandescent lamp (L)

ATOMIC CHEMISTRY EXPERIMENTAL X-RAY APPARATUS—PUPIL PROJECT

CAUTION: X-rays are a type of very penetrating and high energy radiation. Both short exposures to high intensities and long exposures to lower intensities result in dangerous effect on human and animal tissue. This set up *must not* be operated without the lead shield. Stay away from the beam issuing from the hole in the shield! Avoid operating for long periods of time, or do not remain in the immediate vicinity when making a long photographic exposure.

Materials. Ford spark coil or similar induction coil
301 A radio tube, silver coated
sheet lead at least 1/8 in. thick, 3/16 inch better
six volt D.C. power supply
hookup wire and switch
wooden stand and clamps
small sheet aluminum foil
dental film packs (obtainable from local dentist)

Construction. Secure from a radio shop or discarded set a 301 A radio tube or similar high vacuum tube. The 301 A type is obsolete but can still be found. Mount it on a wood block as shown in FIGURE 40 or use a regular socket if one is available.

Make lead cylinders 3½ inches in diameter to fit both ends of the tube. Seal one end of each cylinder with lead by soldering in a disc of this material. Cut a one-inch hole in the longer cylinder for an exit tube.

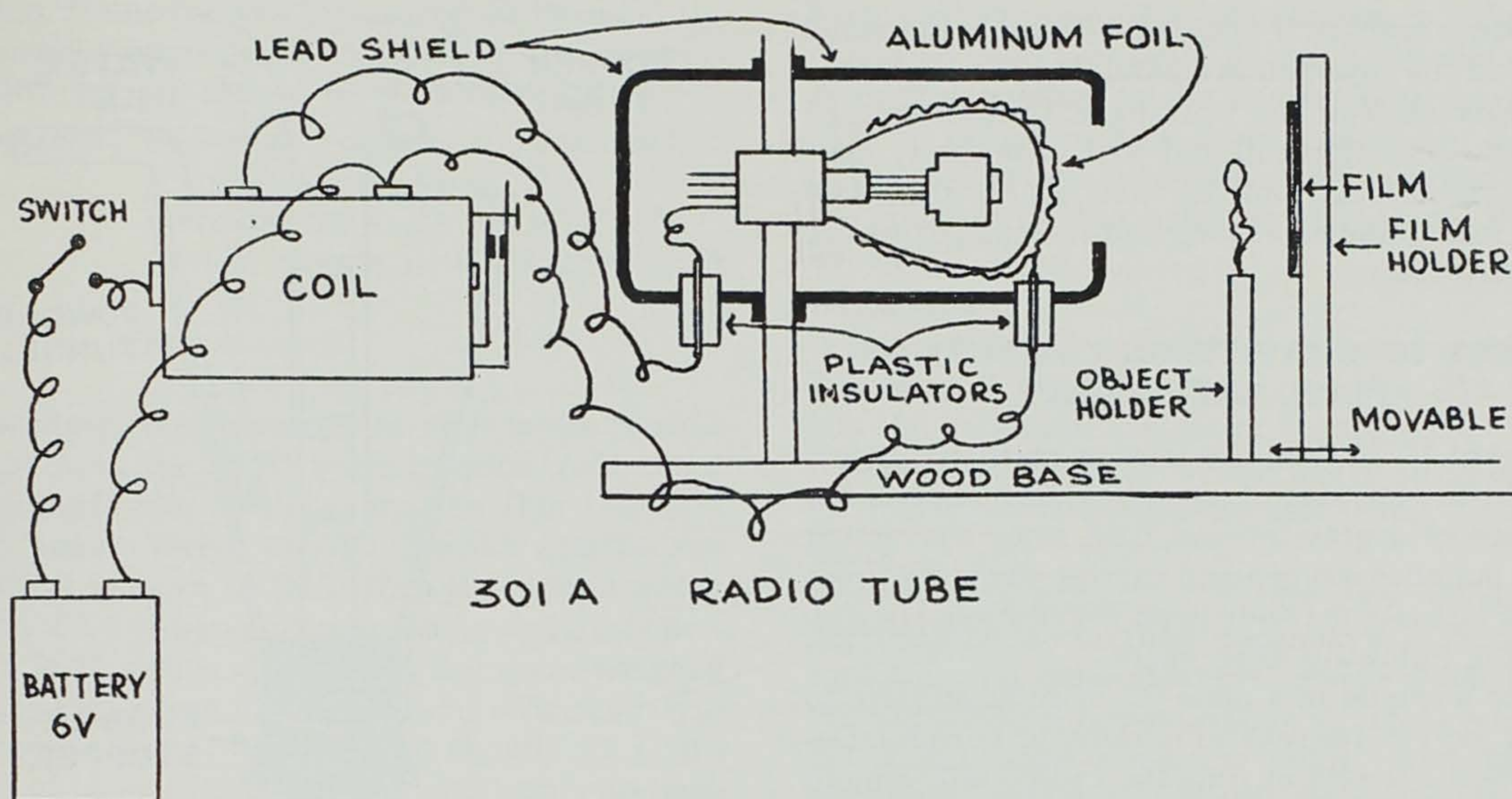


Figure 40

Secure two plastic tubes for insulators and cut holes in the side of each of the lead cylinders to just fit the plastic tubes. Solder a wire to the filament leads (these will be the two heavier ones) of the tube. Run this wire through the plastic tube of the shorter cylinder and attach the cylinder as a cover for the rear end of the tube as shown in the diagram. Attach a second wire to the aluminum foil. Fold the foil around the round end of the glass bulb of the radio tube holding it in place with Scotch tape. Run the other end of the wire inside the longer lead shield, down through the plastic tube, and then attach the lead shield over the front of the tube as shown in the diagram. Connect this wire to the top terminal of the Ford coil nearest the "buzzer," and run another wire from this same contact to the power supply. Run the wire from the back or base of the tube to the other top terminal of the coil. The rear terminal of the coil should be connected to the switch and from the switch to the other side of the power supply.

Directions. Adjust contacts of the coil so that they vibrate evenly producing a good spark between the secondary terminals. Place a dental X-ray film pack on the film holder attaching it with Scotch tape or better by metal clips. Place an object on the object support (a key is good to start with) and move the film stand so the film and object are close together. Make a trial exposure. It is hard to even suggest a length of time for this first exposure as different pieces of equipment vary so much in the intensity of the radiation produced. With a metal object such as a key

an overexposure is better than an underexposure. Start with one of several minutes, but do not remain in the vicinity while waiting. Remove the film to a dark room and develop it according to procedure recommended by the dentist from whom you obtained it or by its manufacturer. If the film is overexposed, cut the next exposure to $\frac{1}{2}$ or $\frac{1}{4}$. If it is underexposed, expose the next film by a power of two, as 2 times, 4 times, etc., until a suitable exposure results. When a suitable exposure has been found, other objects of more subtle X-ray densities may be photographed, e.g., small animals or plants.

Suggested Experiments

Intensity. Make several sheets of aluminum or thin sheet iron and photograph them arranged in echelon fashion. This will give a range of exposures on one photograph. Try plotting exposure times against thickness of absorber. Make a simple densitometer with a light source and a photographic exposure meter of the photocell type and with this estimate the relative densities of the various portions of your film. Plot density of the film against logarithms of the exposures.

Crystals. Mount a crystal of salt or potassium chloride in a small hole in a lead plate and place it in the X-ray beam. Expose a film placed several centimeters away from this crystal for a considerable period. If you are successful, you should secure a Laue diagram of the crystal structure. See reference books on crystal structure.

Wave length. Consult a reference book on X-ray

techniques and see if you can determine the wave length of the rays you are obtaining.

Be careful. You are working with high voltage from the induction coil and with potentially dangerous radiations from the tube. You will be safe if you follow the directions and the suggestions of your instructor.

EFFECT OF GRAVITY ON COMBUSTION—PUPIL EXPERIMENT

Richard M. Sutton, the famous physicist, points out that the burning of a candle depends upon the presence of a gravitational field since convection which supplies oxygen and carries away the products of burning depends upon gravitation. He suggests the following experiment.

Place a candle in a glass jar of sufficient size to supply oxygen for several minutes of burning. Determine the length of time the candle will remain burning in a fixed position as a check on the validity of the experiment.

Drop the jar containing the freshly lighted candle out of a window into a bed of sand or soft earth to cushion the fall and prevent breaking of the glass. The candle will go out during the fall. Explain the physics and chemistry of the phenomenon.

DEMONSTRATING CONDUCTIVITY OF ELECTROLYTES—PUPIL PROJECT

A device that shows the effect of different solutes, or the effect of dilution on the conductivity of solutions can be readily constructed from a funnel and short pieces of glass tubing.¹ The diagram of the apparatus is shown in FIGURE 41.

A light bulb or ammeter can be used to detect the flow of current. A projection meter² works quite well and the whole class can follow the course of events.

Electrolytic conductivity can be shown by allowing tap water to flow through the device. The gap on the central tube should be adjusted so that the lamp does not light. Add a small quantity of an electrolyte. The lamp should light as soon as the salt dissolves. The intensity of the light will decrease as the salt is diluted showing the dependence of conductivity on the concentration of the electrolyte.

¹ H. A. Suter and Lorraine Kaelber, "Apparatus for the Demonstration of Conductivity of Electrolytes," *Journal of Chemical Education*, XXXII (1955), 640.

² Central Scientific Co. Cat. No. 82550 AC-DC Projection Meter P.G. 2565. Use with horizontal, or preferably with vertical overhead slide projector.

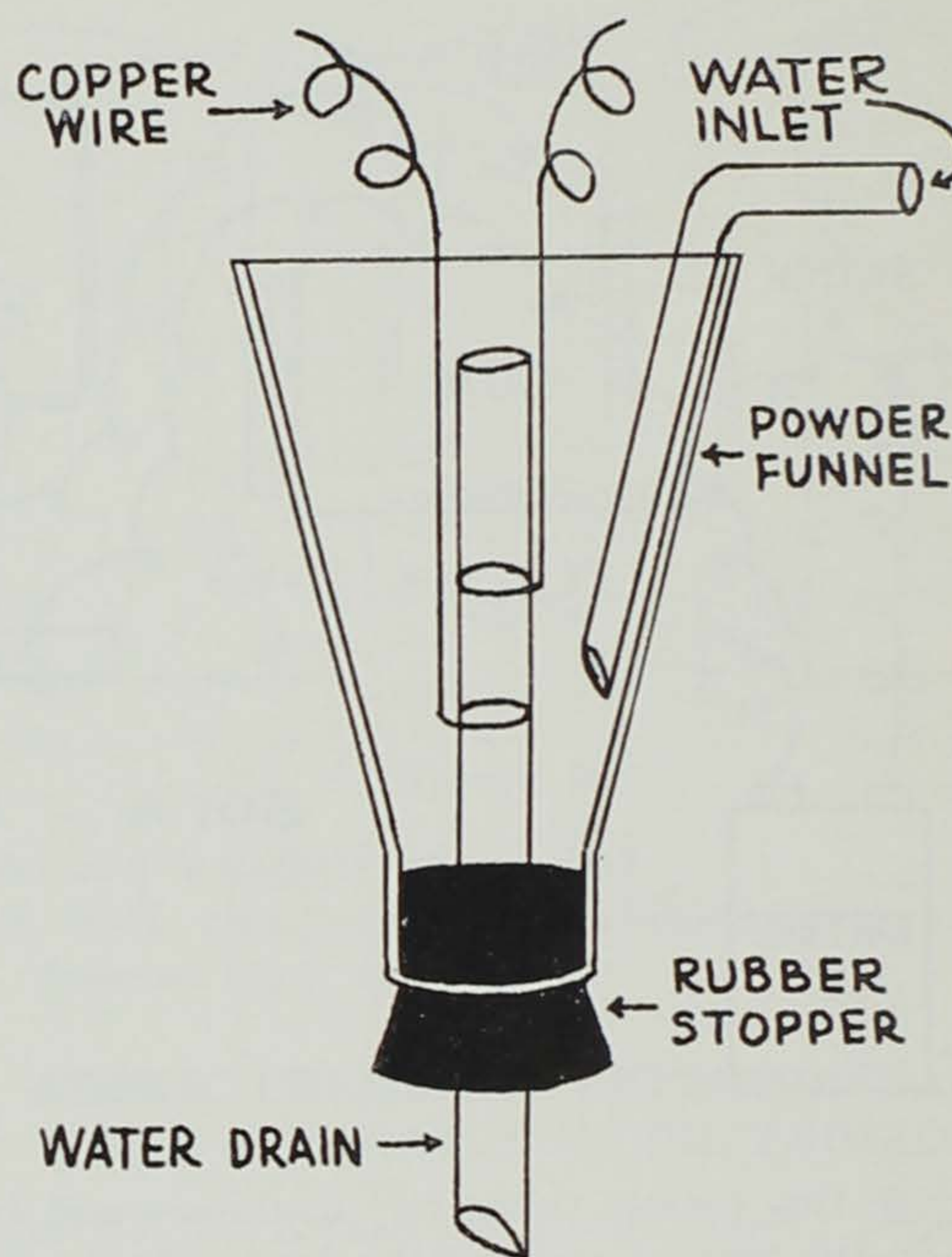


Figure 41

Different electrolytes can be added to the funnel, the reading of the meter or intensity of the light noted, then substance flushed out and a new substance added.

MOLECULAR WEIGHTS BY FREEZING POINT DEPRESSION—PUPIL PROJECT

Introduction

In this experiment you will determine the melting point of a pure solvent; the melting point of a mixture of benzoic acid and the pure solvent; and finally, the melting point of a mixture of unknown and the pure solvent. From the results you should be able to determine the approximate molecular weight of the unknown.

Cyclohexanol with a melting point slightly above room temperature, its excellent solvent properties, and a large cryoscopic constant is almost an ideal solvent for determining the molecular weights of nonvolatile compounds.³ Cyclohexanol is hygroscopic so precautions must be taken to prevent exposure to atmospheric moisture during the experiment. Cyclohexanol has some toxic properties and can be absorbed through

³ This experiment was adapted from an article by Robert Mikulak and Olaf Runquist, "Molecular Weights by Cryoscopy: A General Chemistry Laboratory Experiment," *Journal of Chemical Education*, XXXVII (1961), 557-58.

the skin. If one follows the procedure outlined below and avoids personal contact with cyclohexanol, little hazard should be encountered.

Materials. Sealed vial of pure cyclohexanol
2 empty 6-dram vials
thermometer (-10 to 110°C)
rubber syringe bottle stopper to fit vials
benzoic acid
unknowns for pupils

Procedure. Weigh one of the vials to the nearest 0.01 gram on the triple beam balance. Add about one-half of your cyclohexanol to this vial and weigh again. Weigh enough benzoic acid on the analytical balance so that when it is added to the solvent a 2 to 3 per cent solution by weight is prepared. If 10 grams of cyclohexanol were weighed into the vial, about 0.2 to 0.3 grams of benzoic will be needed. Do not try to weigh exactly 0.2 or 0.3 grams of benzoic acid. It is only necessary to have a weight between 0.2 and 0.3 grams and know this weight to the nearest 0.001 gram. Add the benzoic acid carefully to the vial containing the known amount of cyclohexanol. Benzoic acid is hard to handle without loss. It is best to make a pellet from the crystals before weighing by placing some in a semimicro funnel tube and ramming tight with a glass rod plunger.

Fit the thermometer and the syringe stopper into the vial containing the pure cyclohexanol. Moisten the thermometer with the solvent before attempting to insert the thermometer. Adjust the thermometer until the bulb is within $\frac{1}{2}$ centimeter of the bottom. Place the vial into a beaker of cold water, 15°C . The water level in the beaker should be slightly higher than the solution in the vial. Stir the contents of the vial with the thermometer until a thick mush is formed. Remove the vial and thermometer from the beaker. Hold the vial in your hand to warm slowly while stirring with the thermometer. Take as the melting point the temperature at which the last crystal of cyclohexanol disappears. Repeat the determination at least three times and average the values obtained.

Remove the thermometer and stopper from the vial of pure cyclohexanol. Wipe rubber bulb and thermometer clean and place both in the vial containing the cyclohexanol-benzoic acid solution. Determine the melting point of the solution by the same method outlined above. Repeat at least three times to determine an average melting point.

Weigh the remainder of the pure cyclohexanol into the second vial to the nearest 0.01 gram. Add enough of the unknown to the vial to make a 2 to

3 per cent solution. If your unknown is a liquid, it would be good technique to weigh the vial on the analytical balance to the nearest 0.001 gram and add the unknown with a clean medicine dropper until the desired weight is obtained. Determine the melting of the solution containing the unknown.

Calculations

The cryoscopic constant for cyclohexanol is 39.3. The formula weight of cyclohexanol is 98.2. Calculate the molecular weight of benzoic acid.

Repeat calculations of the above to determine the molecular weight of the unknown.

Notes

Unknowns suitable for this experiment are: naphthalene, acetanilide, dactanoic acid, triphenolcarbinol, and chlorobenzene. See original article for additional suitable unknowns. The cyclohexanol should have a melting point about 24°C . It may be necessary to redistill the cyclohexanol at reduced pressure and collect the material freezing at 24° higher.

COLORMETRIC DETERMINATIONS OF pH— PUPIL PROJECT

Preparation of Buffered Solutions

Walter R. Carmody¹ discusses a method for preparing a set of buffer solutions from pH 2 to pH 12.0 in 0.5 pH units.

Reagent grade, anhydrous boric acid, citric acid (monohydrate)* and tertiary sodium phosphate ($12\text{H}_2\text{O}$) are dissolved into two stock solutions. Solution A consists of 0.200 M boric acid and 0.050 M citric acid while Solution B consists of 0.100 M tertiary sodium phosphate. A table showing the volumes of Solutions A and B required to prepare 200 ml of buffered solution at each pH value is shown below.

pH	Solution A volume (ml)	Solution B volume (ml)
2.0	195	5
2.5	184	16
3.0	176	24
3.5	166	34
4.0	155	45
4.5	144	56
5.0	134	66
5.5	126	74
6.0	118	82
6.5	109	91

¹ Walter R. Carmody, "Easily Prepared Wide Range Buffer Series," *Journal of Chemical Education*, XXXVIII (1961), 559.

* Use only clear crystals free of white effloresced material.

7.0	99	101
7.5	92	108
8.0	85	115
8.5	78	122
9.0	69	131
9.5	60	140
10.0	54	146
10.5	49	151
11.0	44	156
11.5	33	167
12.0	17	183

Volumes of acid Solution A and basic Solution B required for the preparation of 200 ml of solutions in buffer series.

Place 5 ml of each buffered solution in a clean four inch test tube. Add two (or three) drops of universal indicator to each of the solutions and stopper.* Label each color standard in terms of its pH. In place of the universal indicator, individual indicators may be used that cover a specific range. Each indicator can be used for about four pH intervals.

Preparation of the buffered solutions should be done by the instructor or a few able pupils. Preparation of the color standards could be a group project.

Some method is needed to observe and compare the various indicator colors. A support rack capable of holding the ranges selected should be made.

A. Alternate method: (not recommended for best results).

It is possible to prepare solutions at intervals of 1 pH by diluting solutions of 0.1M HCl and 0.1M NaOH tenfold with distilled water. For example, take 5 ml of 0.1M HCl and dilute to 50 ml of recently boiled distilled water. This solution will have a pH 2. Then take 5 ml of the pH 2 solutions

* In case semipermanent standards are desired, it is necessary to sterilize the solutions before sealing. This is done by placing each filled tube in a bath of boiling concentrated CaCl_2 solution until the buffer solution itself boils. Then each tube should be sealed immediately with a paraffined cork, rubber stopper, or better by sealing the glass tube with a glassblowing torch.

and dilute to 50 ml. This resulting solution will have a pH of 3. Repeat until a pH of 6 is reached. Solutions of pH 4, 5, and 6 should be buffered for the best results.

Starting with 0.1M NaOH develop the ranges above pH of 7 in a similar matter for the acid range.

B. Colorimetric determinations of pH.

1. Determination of pH range of an indicator.

Use the 7 pH ranges of the alternate method or preferably the pH ranges of the buffered solution from pH of 2.5 to 6.5, placing 5 ml of each solution in separate clean, dry test tubes. Repeat using methyl violet, methyl red, or bromthymol blue.

2. Determination of the pH of unknown solutions.

Secure an unknown from your instructor. Add 2 drops of universal indicator solution to 5 ml of the solution. Match the color of your unknown with the color standards to determine its pH. Check your results by commercial pH meter or the pH meter from the project section of this outline.

3. pH of a strong and a weak acid.

Add 2 drops of universal indicator to 5 ml portions of 0.1N HCl and of 0.1N acetic acid. Determine the pH of each by colorimetric means. Account for the differences in pH of the acids.

4. Hydrolysis of salts.

Add 2 drops of universal indicator solution to 5 ml portions of 1M solutions of NaCl, NH_4Cl , Na_2CO_3 , AlCl_3 , $\text{NaC}_2\text{H}_3\text{O}_2$. Determine the pH of the various salt solutions.

5. Per cent of ionization of acetic acid.

Determine the pH of 0.1N and 1N acetic acid solutions by colorimetric means. Account for the differences in H_3O^+ concentrations. Compute the H_3O^+ concentration of each solution. From the H_3O^+ concentration determine the per cent of ionization of each solution.

DETERMINATION OF THE FARADAY (ALTERNATE)

Data

Time—start of electrolysis	1:05-00
Time—end of electrolysis	1:21-40
Current during electrolysis	250 ma
Weight of copper electrode before electrolysis	18.093 gms
Weight of copper electrode at end of electrolysis	18.010 gms

Results

Time—seconds	1000
Current—amps	0.250
Coulombs used during electrolysis	250 Q
Weight of copper oxidized	0.083 gms
Equivalent weight of copper (Cu)	31.7 gms
Coulombs/ equivalent weight of copper	96,501 (int. coul. per g. equiv.)

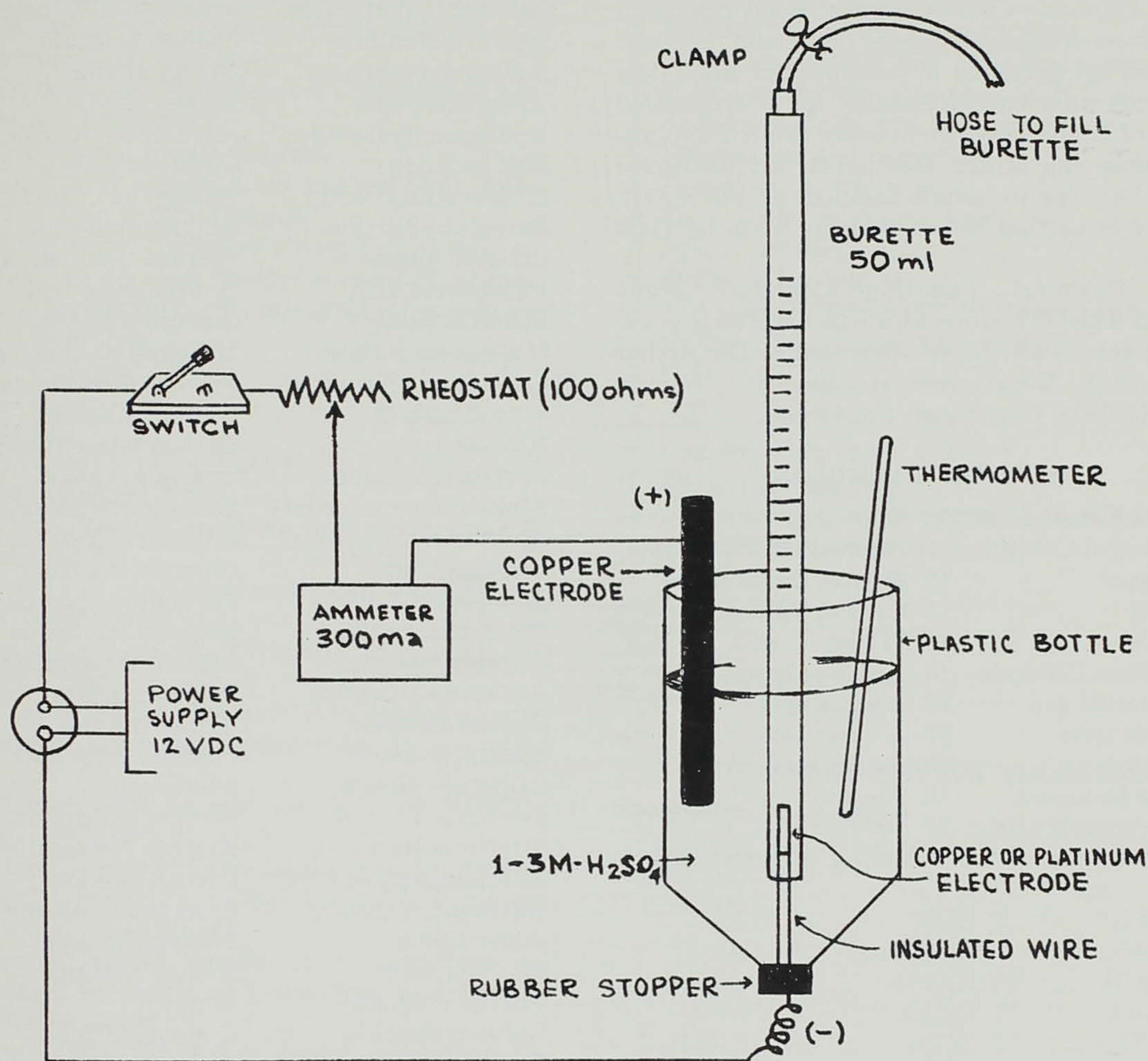


Figure 42

STANDARDIZED EXAMINATIONS AVAILABLE

The best standardized examination available covering the entire chemistry course as usually presented in high schools is the A.C.S.-N.S.T.A. Cooperative Examination in HIGH SCHOOL CHEMISTRY. The current form is that of 1949, but new forms become available about every three years, and older forms are usually available. These examinations are five-choice, objective-type tests of about 100 items taking from 75 to 90 minutes depending upon the form used. Answers are made on printed answer sheets and are either hand or machine scorable. Nationally standardized scores are available for comparison of scores in terms of percentiles. For information and orders, address Dr. Theodore A. Ashford, Chairman, Examinations Committee (A.C.S.), University of South Florida,

Tampa, Florida. Their tests serve admirably for a final examination at the end of the course.

For tests to be used at intervals during the course, several sets are available from publishers to accompany the text adopted. An example is the Chemistry for Progress Tests by Young and Petty which is intended to be used in conjunction with the above authors' text of that name. These are available from Prentice-Hall, Inc., Englewood Cliffs, New Jersey. With some selection of areas covered, many of these individual unit tests may be used with the Iowa Plan Outline. A similar set by Charles Dull and Joseph Castka is available from Henry Holt and Company, New York, New York.

A set of tests, including a final test, is available from the Chemical Bond Approach Committee. These tests are available from Dr. Laurence E.

Strong, Editor in Chief, Chemical Bond Approach Committee, Earlham College, Richmond, Indiana. The tests are excellent in makeup; but since they cover the material presented in a specialized course of study, they are of more limited applicability than the others mentioned. An instructor would be wise to secure samples of the C.B.A. tests for inspection before ordering them for class use.

The Chemical Education Material Study (CHEM STUDY) group also will have tests available. Inquiries should be addressed to Dr. Arthur J. Campbell, Department of Chemistry, Harvey Mudd College, Claremont, California.

WRITING FORMULAS

Write formulas for the following:

- | | |
|---------------------|------------------------|
| 1. Calcium Chloride | 11. Ammonium Hydroxide |
| 2. Copper | 12. Barium Sulfate |
| 3. Iron | 13. Carbon Monoxide |
| 4. Arsenic | 14. Water |
| 5. Sodium Chloride | 15. A strong base |
| 6. Chlorine gas | 16. A weak acid |
| 7. Hydrogen | 17. A weak base |
| 8. Nitrogen | 18. An acid salt |
| 9. The Halogens | 19. An acid anhydride |
| 10. Phosphoric acid | 20. An oxide |

Rewrite any of the following incorrect formulas:

- | | | |
|-----------------------------|---------------------------------|---|
| 1. NaCl_2 | 8. HOH | 15. H_2O_2 |
| 2. CaSO_4 | 9. $\text{Ca}_2(\text{PO}_4)_3$ | 16. AgCrO_4 |
| 3. MgOH_2 | 10. Na_3PO_4 | 17. Sb_3S_2 |
| 4. H_3SO_4 | 11. CuSO_4 | 18. CdNO_3 |
| 5. HNO_3 | 12. Sn_2S_4 | 19. $\text{NH}_4\text{C}_2\text{H}_3\text{C}_2$ |
| 6. K_2ClO_3 | 13. SO_2 | 20. IK |
| 7. Na_2CO_3 | 14. ZnO | 21. AgCN |

Make compounds out of the following groups of elements:

Example

carbon	hydrogen	potassium
hydrogen	oxygen	oxygen
oxygen	nitrogen	hydrogen
H_2CO_3		

oxygen	sulfur	sodium
hydrogen	oxygen	oxygen
sulfur		nitrogen

Write the formulas for the following compounds:

Magnesium phosphate	Aluminum chloride
Potassium periodate	Potassium permanganate
Plumbus acetate	Ammonium carbonate
Sodium sulfide	Nickel sulfide
Ferric nitrate	Sodium thiocyanate

Calcium hypochlorite	Magnesium cyanate
Mercurous sulfate	Sodium bromate
Sodium bicarbonate	Nickel nitrite
Arsenious oxide	Arsenic acid
Hydrogen hydroxide	Mercurous chloride
Zinc bromide	Cupric sulfate
Hydrosulfuric acid	Aluminum cyanide
Barium hydroxide	Sodium phosphide
Cobaltic nitrate	Ammonium hydroxide
Phosphorus acid	Hydrofluoric acid
Bismuth oxide	Stannic chloride
Manganese sulfate	Mercuric phosphate
Chromium hydroxide	Ferrous dichromate
Silicon carbide	Sodium arsenite
Nitrous acid	Carbon tetrachloride
Carbon monoxide	Cuprous cyanide
Ammonium molybdate	Strontium hydroxide
Aluminum carbonate	Lithium silicate
Chromic acid	Plumbic chloride
Manganese dioxide	Phosphrous pentoxide
Potassium iodide	Magnesium nitride
Stannous nitrate	Aluminum selenide
Antimony trisulfide	Ferric sulphate
Carbon disulfide	Sulfur trioxide
Sodium chlorate	Calcium carbonate
Cadmium nitrate	Acetic acid
Ferric oxide	Sodium hydroxide
Stannic sulfate	Arsenic Pentasulfide
Aluminum phosphate	Carbon dioxide
Ammonium acetate	Ammonium nitrate
Silver sulfide	Phosphoric acid
Zinc arsenate	Sodium bromide
Titanium dioxide	Water
Calcium fluoride	Cuprous nitrate
Sulfurous acid	Ammonia

CHEMISTRY REVIEW QUESTIONS

- List the physical properties of H_2 , of O_2 .
- Which of these are found in the free state?

Oxygen	Fe	H_2O
Cu	Na	Sulfur
H_2	Pb	He
- What causes iron to rust?
- Air is a (a) mixture, (b) compound, (c) element, (d) none of these.
- When coal in a furnace burns, (a) the gas from the coal unites with the carbon to give off heat, (b) the oxygen of the air combines with the coal, gas, and carbon to give off heat and oxides of carbon, (c) black visible smoke must be one of the products, (d) oxidation has taken place.
- It is estimated that the per cent of oxygen in the earth's crust is (a) 21%, (b) 88.8%, (c) 50%, (d) 78%.

- Four-fifths of the air is (a) oxygen, (b) hydrogen, (c) carbon dioxide, (d) nitrogen, (e) water vapor, (f) rare gases.
- List the important chemical properties of oxygen and hydrogen.
- Describe the Frasch process.
- Name three organic acids and tell their uses.
- What is an acid anhydride?
- Of what use is Avagadro's law?
- What is a physical change? a chemical change?
- How are the following prepared commercially?
 - Hydrogen
 - Oxygen
 - H_2SO_4
 - HCl
- When salt is put in H_2O , the salt is the _____ and the water is the _____.
- In the previous solution, pure water may be obtained by _____ the solution.
- A solution containing much acid and little water is said to be _____.
- An (electrovalent, covalent) compound will ionize readily.
- What is the formula for the lime used on soil?
- Litmus paper and phenolphthalein are called _____.

- Which is NOT an acid (a) HCl , (b) NaHCO_3 , (c) HNO_3 , (d) H_2CO_3 ?
- Which does NOT increase the speed of dissolving (a) heat, (b) cooling, (c) pulverizing, (d) mechanical shaking?
- The molecular weight of a compound expressed in grams and dissolved in a liter of water is called _____.
- The density of H_2O is (more, less) than the density of mercury.
- Air is an (a) element, (b) mixture, (c) compound, (d) all of these.
- Which will produce H_2 gas (a) $\text{Zn} + \text{H}_2\text{SO}_4$, (b) $\text{AlCl}_3 + \text{H}_2\text{O}$, (c) $\text{NaOH} + \text{HCl}$, (d) $\text{Na} + \text{H}_2\text{O}$?
- Name the largest class of inorganic compounds.
- Write out the chemical name these symbols stand for:

C	Cu	Hg
K	Mg	Pb
Na	N	Au
Fe	P	Al

- Draw the atomic structure of these atoms.
Chlorine Carbon Potassium

EQUATIONS

Balancing equations; complete and balance the following equations:

- Aluminum + sulfur \rightarrow
- $\text{Ca}(\text{OH})_2$ + sodium phosphate \rightarrow
- Mercurous nitrate and barium chloride \rightarrow
- Hydrogen sulfide + oxygen $\rightarrow \text{H}_2\text{O} + \text{SO}_2$
- Sodium chloride and sulfuric acid \rightarrow
- Phosphorus and oxygen \rightarrow
- Potassium + water \rightarrow
- $\text{Fe} + \text{steam} \rightarrow \text{FeO} + \text{Fe}_2\text{O}_3 + \text{_____}$
- $\text{NH}_3 + \text{HOH} \rightarrow$
- Sulfur + oxygen \rightarrow
- The reaction between sulfuric acid and NaNO_3
- $\text{KOH} + \text{_____} \rightarrow \text{KCl} + \text{H}_2\text{O}$
- $\text{H}_2\text{O} + \text{CO}_2 \rightarrow \text{_____}$
- $\text{CaCO}_3 + \text{_____} \xrightarrow{\Delta} \text{CO}_2 + \text{_____} + \text{_____}$
- Copper oxide (hot) + $\text{H}_2 \rightarrow$
- $\text{Zn} + \text{a strong acid} \rightarrow$
- Commercial prop of hydrogen in which CO is given off as a by-product
- $\text{Mg (hot)} + \text{oxygen} \rightarrow$
- $\text{KBr} + \text{Cl}_2 \rightarrow$
- $\text{HgO} \xrightarrow{\Delta}$
- $(\text{NH}_4)_2\text{CO}_3 \xrightarrow{\Delta}$

CHEMISTRY REVIEW PROBLEMS

Percentage Composition

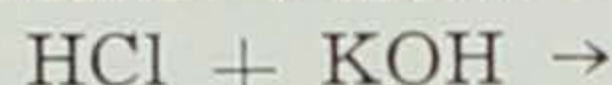
Calculate the per cent of each element in the following formulas:

- | | |
|--------------------------------------|-----------------|
| 1. $\text{Al}_2(\text{SO}_4)_3$ | %Al |
| | %S |
| | %O |
| 2. CuSO_4 | %Cu |
| | %S |
| | %O |
| 3. $\text{HC}_2\text{H}_3\text{O}_2$ | %H |
| | %C |
| | %O |
| 4. AgCN | %Ag |
| | %C |
| | %N |
| 5. NaHCO_3 | %Na |
| | %H |
| | %C |
| | %O |
| 6. H_2CO_3 | %H |
| | %C |
| | %O |
| | % CO_2 |

Normality and Molarity

- 3 moles of HCl in 1 liter of solution make a solution _____ Molar _____ Normal.
- 2 moles of H_3PO_4 in 0.5 liter of solution make a solution _____ Molar _____ Normal.

3. In this reaction, how many moles of acid are used to neutralize 3 moles of base?



4. How many ml of 6N base are required to neutralize 20 ml of 3N acid?

5. $\text{CaCO}_3 + \text{HCl} \rightarrow$

How many liters of CO_2 are produced in this reaction if 0.5 liter of 3N HCl and an excess of calcium carbonate are used?

Molecular Weight Problems

1. What weight KClO_3 must be heated to make 12 liters of oxygen at S.T.P.?

2. How much CO_2 in grams will 20 g CaCO_3 make when reacted with 7 grams of HCl?

3. How many grams of silver sulfide will be precipitated when 100 liters of pure H_2S at S.T.P. are run into a saturated silver nitrate solution?

4. How many liters of CO_2 at S.T.P. will be produced when 1 kilogram of limestone is decomposed with heat?

5. Natural gas, methane, (CH_4) burns in air to form water and carbon dioxide. How many ml of water are produced when 500 liters of CH_4 at S.T.P. are burned?

CHEMISTRY APPARATUS

Section A—Pupil Unit Items (One Set-up for Each Two Pupils)

Purchase Guide Number NDEA	Quantity	Description
	2	Apron, Rubber Large
0370	2	Beakers, 150 ml Pyrex
0370	2	Beakers, 250 ml Pyrex
0370	2	Beakers, 400 ml Pyrex
0540	2	Bottles, 1 liter, Narrow Mouth
0540		Bottles 4 oz., Wide Mouth Lot of 12
0540		Bottles 8 oz., Wide Mouth Lot of 6
		Brushes
0595	1	Burette, 50 ml, Mohr
0605	1	Burner, Bunsen-adjustable
0880	2	Clamps, Burette, single
0880	2	Clamps, Condenser
0880	2	Clamp Holder, Right Angle
0880	2	Clamp, Test Tube
0880	2	Clamp Tubing, Mohr's pinch cock
1040	1	Condenser, Straight Jacket 400 mm
1150	1	Crucible, No. 1
1140	1	Crucible Cover, No. 1
1240	1	Dish, Evaporating, No. 0
	2	Drying Tubes, CaCl_2
1520	1	Flask, Boiling 500 ml
	1	Flask, Distilling 250 ml
	1	Flask, Round Bottom 200 ml
	1	Flask, Erlenmeyer 50 ml
1510	2	Flask, Erlenmeyer 250 ml Wide Mouth
1595		Forceps, 5"
1630	1	Funnel, 65 mm, Long Stem
1180	1	Graduate, 100 ml
1180	1	Graduate Guard
2085	6	Glass Plates, Sq. 3 × 3 in.
3240	6	Dropper
		Matches, Safety
2935		Mortar and Pestle No. 00

3235	1	Pipette, Measuring 10 ml Graduated in 1/10 ml
		Paper, Filter
		Paper, Litmus (Red and Blue)
4395		Trough, pneumatic
4020		Ring Stand, 3 Ring
3910		Spoon, Deflagrating
4030		Support, Test Tube, Wood, 12 place
4125	12	Test Tubes, 15 × 150 mm
4125	6	Test Tubes, 25 × 150 mm
4175	1	Thermometer, 150° C.
1640		Thistle Tube
		Tongs, Crucible, Stainless Steel
		Towel
4370		Triangle, 2½"
		Tube, Pyrex, 18 × 150 mm
0365		Watch Glasses, assorted to fit beakers
0605		Wing Top, 7/16", to fit burners
		Wire Gauze, Asbestos Center 5 × 5 in.

SUGGESTED LIST OF CHEMICALS

This is a list of chemicals such as are often used in general chemical laboratories. The list is necessarily incomplete as no attempt has been made to include chemicals required for those experiments contained in other publications and listed as references. It does include all chemicals for experiments described in this outline. The amounts of each chemical needed will have to be estimated by each instructor on the basis of the number of pupil units to be serviced in the laboratory. The suggested amounts now on the list are only sufficient to take care of class demonstrations.

Acids and Other Inorganic Chemicals

Quantity	
5 lbs. scb*	Acid, acetic, glacial CP
4 × 6 lbs. scb*	Acid, hydrochloric CP
7 lbs. scb*	Acid, nitric CP 1.42 sp. gr.
9 lbs. scb*	Acid, sulfuric CP 1.84 sp. gr.
1 lb.	Aluminum metal, powder
1 lb.	Aluminum nitrate (or chloride) CP
1 lb.	Aluminum sulfate CP
5 lbs.	Ammonium chloride CP crystals
2 × 4 lb. scb*	Ammonium hydroxide CP
1 lb.	Ammonium thiocyanate CP crystals
2 × ¼ lb.	Antimony trichloride CP crystals
1 lb.	Barium chloride CP
2 × ½ lb.	Barium oxide CP
1 lb.	Bromine
1 lb.	Cadmium chloride CP

1 sq. foot	Cadmium metal sheet
1 lb.	Calcium acetate CP
1 lb.	Calcium carbide
1 lb.	Calcium carbonate
5 × 1 lb.	Calcium chloride anhydrous 8 mesh (free of powder)
1 lb.	Calcium chloride CP crystals
1 lb.	Calcium hydroxide CP
4 × ¼ lb.	Calcium oxide (must be anhydrous)
1 lb.	Calcium sulfate CP
¼ lb.	Cesium chloride CP
¼ lb.	Chromium chloride
¼ lb.	Cobalt chloride or cobalt nitrate CP
8 × 2 oz. spools	Copper wire, bare
1 lb.	Cupric oxide wire
5 lbs.	Cupric sulfate CP
1 lb.	Ferric chloride CP
1 lb.	Ferric oxide
1 lb.	Ferrous sulfate CP
1 case (4 bottles)	Hydrochloric acid
16 oz.	Hydrogen peroxide, 30 per cent
1 lb.	Iodine CP
1 coil	Iron wire (picture wire)
1 lb.	Lead acetate CP
1 lb.	Lead nitrate
1 sq. foot	Lead sheet, thin
¼ lb.	Lithium chloride
4 oz.	Magnesium ribbon
5 lbs.	Manganese dioxide
½ lb.	Mercuric oxide
5 lbs.	Mercury
1 lb.	Nickel chloride or nitrate

* Scb indicates screw cap bottle. These bottles are now being shipped in foamed-plastic cases of four bottles each. Case orders save money and extra expense in handling as well.

1 sq. foot	Nickel sheet
1 case	Nitric acid
1 lb.	Phosphorous, red
1 lb.	Phosphorous, white
1 foot	Platinum wire No. 20 or No. 22
1 lb.	Potassium bitartrate CP for buffer use
5 lbs.	Potassium chloride CP
1 lb.	Potassium chromate CP
1 lb.	Potassium hydroxide CP
½ lb.	Potassium iodate CP
2 × 1 lb.	Potassium iodide CP
1 lb.	Potassium permanganate CP
¼ lb.	Rubidium chloride
1 lb.	Silicon dioxide
4 sq. inches	Silver foil
1 lb.	Silver nitrate CP or reagent
5 lbs.	Sodium acetate CP crystals
1 lb.	Sodium bicarbonate CP
1 lb.	Sodium bromide CP
5 lbs.	Sodium chloride CP crystals
1 gal.	Sodium hydroxide 50 per cent solution CP
1 lb.	Sodium metal CP (caution, after opening container, store under light oil or kerosene. Do not allow contact with water or halogenated organic compounds.**
¼ lb.	Sodium peroxide*** CP
1 gal.	Sodium silicate solution
1 lb.	Sodium sulfate CP
5 lbs.	Sodium thiosulfate CP
1 lb.	Stannous chloride CP dihydrate
¼ lb.	Strontium chloride CP
5 lbs.	Sulfur flowers, purified
1 lb.	Wood's metal alloy
5 lbs.	Zinc, mossy or small shot
1 sq. foot	Zinc sheet
1 lb.	Zinc sulfate or nitrate CP

Organic Chemicals and Miscellaneous

5 lbs. scb*	Acid, acetic (see inorganic acids)
1 lb. scb*	Acid, formic CP
½ lb.	Acid, oleic
1 lb.	Acid, phthalic
1 lb.	Acid, pyrogalllic
¼ lb. or smallest pkg.	Adipyl chloride
1 pkg.	Alconox or Dreft, Tide, Joy, or other detergents
2 oz.	Aspirin tablets
5 lbs.	Benzene

* Screw cap bottle.

** Explosion danger.

*** Hazardous chemical, fire danger.

½ lb.	Bone black powder
½ lb.	Camphor, gum
5 lbs.	Carbon tetrachloride
8 oz.	Cotton, absorbent or plugging
2 oz.	Cotton, natural
1 lb.	EDTA, ethylene, diamine, tetraacetic acid
2 lbs. or normal size container****	Ethyl acetate
2 gals.	Ethyl alcohol
2 oz.	Ferroxyl reagent
1 pt.	Formaldehyde solution 15 per cent
1 cyl.	Freon small cylinder (Virginia Chemical Company "can-o-gas")
1 lb. scb*	Glycerine CP
½ lb. or smallest container	Hexamethylene diamine
10 g.	Indicator, universal, Hartman Leddon Company
1 gal.	Methyl alcohol
¼ lb.	Methyl orange
25 g.	Methyl red (water soluble)
1 pt.	Mineral oil (Nujol)
1 lb.	Naphthalene
	Paper, litmus blue, red
1 lb.	Paraffin
5 lbs.	Phenol
¼ lb.	Phenolphthalein indicator
1 lb.	Cation exchange resin, Dowex 50, Dow Chemical Company, Midland, Michigan
1 lb.	Salicylic acid
¼ lb.	Starch, soluble for indicator use
2 lbs.	Sugar (sucrose)
25 g.	Thymolphthalein, indicator (water soluble)
2 lbs.	Urea

SUGGESTIONS FOR PURCHASING

Apparatus, Equipment, Glassware

Experience has shown that it is usually wiser to purchase apparatus on a bid basis where the total order is over one hundred dollars. Suppliers should be asked to bid "item by item." While many houses prefer to bid on an "all or nothing" basis, accepting such bids does not allow the purchaser to make a final selection of individual items and usually the vaunted extra savings on such bids are more illusory than real. Where the final selection of items is at the discretion of the person for whom the order is drawn rather than a pur-

**** See chemical catalog.

chasing agent or board, item by item bidding will, over a period of time, result in greater savings and acquiring of better merchandise.

Whether bids should be requested on an f.o.b. shipping point or on a cartage paid or delivery basis depends largely on the local situation. When working with a restricted or a fixed budget, having items priced on a delivery basis allows closer figuring. However, if bids are requested on a delivered basis, they should be requested several weeks earlier as it takes the jobbers longer to figure bid prices on this basis.

It is not wise to attempt to purchase all items of apparatus from a single supplier. While some supply houses cater especially to high schools and handle nearly all items needed, university and industrial suppliers are equally willing to quote on high school orders often at favorable prices. This is especially true of electrical supplies. Some manufacturers will sell directly to an institutional consumer, others sell only through supply houses, or even through a few franchised distributors. Such situations can only be determined by correspondence. In any case, it is worthwhile to secure detailed company literature describing a piece of equipment before final selection is made regardless of where the final purchase is lodged.

In the case of expensive or complicated equipment such as spectrophotometers, calculators, balances, etc., the possibility of a maintenance contract with the manufacturer should be investigated. Periodic checking and repair by factory representatives under a maintenance contract is usually cheap in the long run as it results in better satisfaction and longer life for the equipment.

Instrumentation is developing rapidly at the present time. Before purchasing an instrument, it should be ascertained whether the item contemplated is the latest model and whether radical improvements in function or design are in process which will render the current model obsolete. Instruments listed in the catalogs are often superseded by later models described in pamphlets. Unless a careful filing system is maintained, it is easy to lose track of these improvements.

Each year more instruments are being imported from England, Scandinavia, Japan, Europe, and the Soviet Union. Some of these items are excellent and of better design than those produced currently in the United States. Others are definitely less desirable, even though cheaper, because of design or construction. The factors of service and repair should be investigated. Where a choice must be made between an imported and a domestic in-

strument, actual inspection and trial of the instrument is suggested.

When selecting an instrument from among several with similar functions, a number of criteria should be considered in making a final decision. These are: range of measurements, sensitivity, versatility of function, accessories available, simplicity of operation, stability of reading, sensitivity to line voltage variations if electrically operated, probability of long time accuracy, ruggedness under routine or pupil use, reliability of the manufacturer, similarity to current industrial instruments, and cost in relation to the function performed. Often it is better to buy two simple instruments of moderate cost covering two independent but related measurements rather than a single more expensive instrument capable of both measurements. This is because the independent instruments can be used simultaneously by two class groups with more efficient use of class and instructor's time.

For teaching purposes, an instrument designed so that its construction and functioning is evident or easily ascertained is better for class use than one which is merely a "black box" to the pupils. This is especially true where the instrument is to be used largely for demonstrations as the pupils get less chance to learn its functioning. A home-made or "bread board" setup, if not too complicated, is often a better teaching tool than a commercial instrument. If the pupils can enter into the design, assembly, or even construction of a working instrument or demonstration, so much the better. For this reason, directions for several simple instruments are included in the "project" section of this guide.

Regarding the purchase of glassware, several recent changes in the market should be mentioned. For a number of years Pyrex laboratory ware has been standard for articles such as beakers, flasks, etc., which must stand direct heat. Recently, Kimax glass has come on the market which is priced item for item almost, if not identically, the same as Pyrex. These two competing lines appear to be of equal quality, chemical and thermal resistance, and weight. A few other more limited lines of American-made glassware are available and are competitive on certain items. Within the last few years, imported "thermal resistant" glass has appeared on the United States market, coming from Italy and central Europe. This seems to have good resistance to chemicals and to thermal shock, but the items tend to be lighter in weight, thin walled, and hence more subject to breakage. Before exten-

sive purchases of laboratory glassware are made on a bid basis, the quality should be investigated. If the items quoted are not standard brands, samples should be obtained for inspection and trial. For items such as graduated cylinders, funnels, etc., which are never heated, soda lime or soft glass is quite satisfactory and often can be purchased instead of borosilicate glass at a considerable saving.

Chemicals

Chemicals to be purchased in lots of fifty dollars or more should be sought on a bid basis as worthwhile savings can result. When this is done, however, the grade of chemical desired must be carefully specified for each item. The so-called C.P. (chemically pure) or more recently A.C.S. grade chemicals are those which pass the minimum requirements for reagent grade chemicals set up by a committee of the American Chemical Society. Some manufacturers have a quality line of "reagent" or "Analyzed" grade chemicals which at least equal, and in many cases exceed, the A.C.S. grade in purity. U.S.P. chemicals meet the minimum requirements of the United States Pharmacopoeia for drug use. These chemicals are highly purified in terms of contaminants which would be objectionable when taken into the body but may contain larger amounts of other chemical reactions. U.S.P. chemicals are not advisable for use in analytical work but are often quite satisfactory for demonstrations, etc. Technical or "practical" grade chemicals, especially solvents and organic chemicals, are obtainable at considerable savings. Where their suitability for the specific use is known, they may be substituted for the much more expensive reagent grades. They should not be substituted indiscriminately.

In general, in smaller schools where the stock of chemicals is limited, purchase of the best grades is wise, for a single bottle of reagent grade material is suitable for all purposes. If finances and storage space are adequate, savings can be effected by ordering cheaper grades of some chemicals in larger quantities and stocking only a pound or two of reagent grade for more critical uses. Examples of this are roll sulfur for demonstrations, and "blue vitriol" technical $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ for making batteries.

It is probably better to order acids and alkalis in A.C.S. or reagent grade since the use of these is so varied and the stocking of two grades is cumbersome. The possible exception is sulfuric acid. Furthermore, pupils are not critical and often are unaware of the differences in quality. They are likely to use the best grade for technical purposes and vice versa.

Sodium hydroxide can be obtained as solid pellets, but even these are messy to handle and stock. If stored for some time, there is a tendency to cake and to pick up carbon dioxide from the air. Purchase of this chemical as reagent grade 50 per cent aqueous solution is recommended. Since sodium carbonate is almost insoluble in the 50 per cent solution, the stock remains carbonate free and maintenance and dispensing is simplified.

Although considerable saving can be realized by buying chemicals in larger containers, there are several chemicals best purchased in one pound or quarter pound sizes because of their tendency to absorb moisture after the seal is broken. Among these are CaCl_2 anhydrous (1 pound bottle recommended), CaO , ZnCl_2 and other zinc salts, SbCl_3 , P_2O_5 , PCl_3 , PCl_5 , and SnCl_2 . See Lange's Handbook of Chemistry, Handbook Publishers, Inc., Sandusky, Ohio, for precautions for the storage of peculiar and hazardous chemicals.

Local Sources of Materials and Chemicals

It is best to order chemicals needed for a year from regular chemical supply houses in order to secure the purity and quantities necessary. A few chemicals may be obtained locally in even the smaller towns in U.S.P. or better grades often at a considerable saving. These are listed under "A." Under "B" are listed chemicals usually obtainable locally only in less pure or technical grades with their possible sources. This should be considered an emergency list with a few exceptions such as quick lime, sodium hypochlorite, salt, etc. Not specifically listed are many U.S.P. chemicals having common or drug use which can often be obtained from local pharmacists. This supply, however, is often limited in quantity and should not be depended upon for class use unless prior arrangement has been made. Some local pharmacists are willing to bid on entire lists.

List A—U.S.P. chemicals obtainable at grocery stores.

Aspirin, acetyl salicylic acid
 Borax, $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10 \text{H}_2\text{O}$
 Cotton, absorbent
 Glycerine
 Hydrogen peroxide, H_2O_2 , 3 per cent
 Sodium bicarbonate, baking soda, NaHCO_3
 Sodium hydroxide, lye, NaOH . Check purity on label
 Sucrose, cane sugar, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$
 Sulfur

List B—Technical grades usually.

Acetic acid, dilute—see vinegar
 Alum, sodium aluminum sulfate hydrate
 Ammonia, household ammonia, NH_3 solution
 (Note, often has detergent or fine abrasive added)
 Ammonium chloride, sal ammoniac, NH_4Cl
 Ammonium nitrate, NH_4NO_3
 Calcium carbonate, CaCO_3
 Calcium chloride, CaCl_2 (anhydrous or hydrate)
 Calcium hydroxide, $\text{Ca}(\text{OH})_2$, slaked lime
 Calcium hypochlorite, CaClOCl , chlorinated lime
 Calcium oxide, CaO , quick lime
 Calcium polysulfide, CaS_x , lime-sulfur
 Carbon tetrachloride, CCl_4 Carbona
 Formaldehyde, 15 per cent solution
 Glycerine
 Grain alcohol, denatured alcohol
 (Often not satisfactory for chemical purposes, check formula
 against list of denaturants in *Handbook of Chemistry*)
 Grain alcohol, Vodka—dilute, expensive
 Hydrocarbon mixtures, heavy, Nujol
 Hydrocarbon mixtures, light, high test gas
 Hydrocarbon mixtures, medium, Stoddard solvent
 Hydrocarbon mixtures, solid, paraffin
 Hydrocarbons, aromatic with side chain, turpentine
 Isopropyl alcohol, rubbing alcohol
 Manganese dioxide, MnO
 (Usually mixed with ZnCl_2 , NH_4Cl , etc.)
 Methyl alcohol, wood alcohol, antifreeze
 (Antifreeze form usually contains additives)
 Naphthalene, moth balls
 Paradichlorobenzene, moth crystals
 Phenol, carbolic acid
 Potassium bromide, KBr
 Silicon dioxide, sand
 Sodium carbonate, hydrate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, washing soda
 Sodium chloride, NaCl , water softener salt
 (This does not contain KI as does most household salt)
 Sodium hydroxide, NaOH , lye
 Sodium sulfite, anhydrous Na_2SO_3
 Sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$, hypo
 Sulfuric acid, H_2SO_4
 Vinegar, distilled, acetic acid 3 per cent

grocery

grocery

hardware

fertilizer plant

fertilizer plant

highway department

lumber yard

hardware

lumber yard

hardware

grocery

pharmacy, hardware

grocery

hardware

liquor store

grocery, pharmacy

filling station

dry cleaner

grocery

hardware

pharmacy, grocery

old dry cells

hardware, filling station

grocery

pharmacy, grocery

pharmacy

photographer

lumber yard

grocery

grocery

grocery

photographer

photographer

battery service, cream stations

grocery

EQUIPMENT FROM LOCAL SOURCES

The extent and variety of materials and equipment available from local sources other than regular scientific suppliers depends upon the local situation but even more upon the ingenuity of the instructor. The following suggestions are intended to be representative rather than complete.

Variety stores are often good sources of such items as rubber balloons, bubble producing solution, plastic clay for models, plastic foam balls, small tools, wire, etc.

State Surplus Materials distribution depots may contain items of scientific value depending upon what is available for distribution at the time. Such items as tools, metals, electrical wire, bare wire, electrical goods and instruments, optical instruments, paint, mechanical items adaptable to scientific use are not uncommon. Such items are available to state institutions, schools, etc., at cost of handling. In Iowa, the State Surplus Materials depot is located on the State Fairgrounds in Des Moines. To secure such equipment, a person must have official authorization from his school. A list of distribution agencies in other states may be obtained from the Department of Health, Education, and Welfare, Washington, D.C.

Radio houses dealing with war surplus equipment are good sources of electrical components. One must be careful, however, to determine that the items purchased have specifications that fit the needs closely enough to warrant the purchase. Buying an item just because it is cheap is false economy. Names and addresses of firms dealing in surplus electrical equipment may be found in the advertising pages of various magazines for radio amateurs.

Local radio and TV repairmen are usually willing to give junk sets to schools for the parts which they contain. The 6.3 v. windings of radio power transformers (with high voltage leads well isolat-

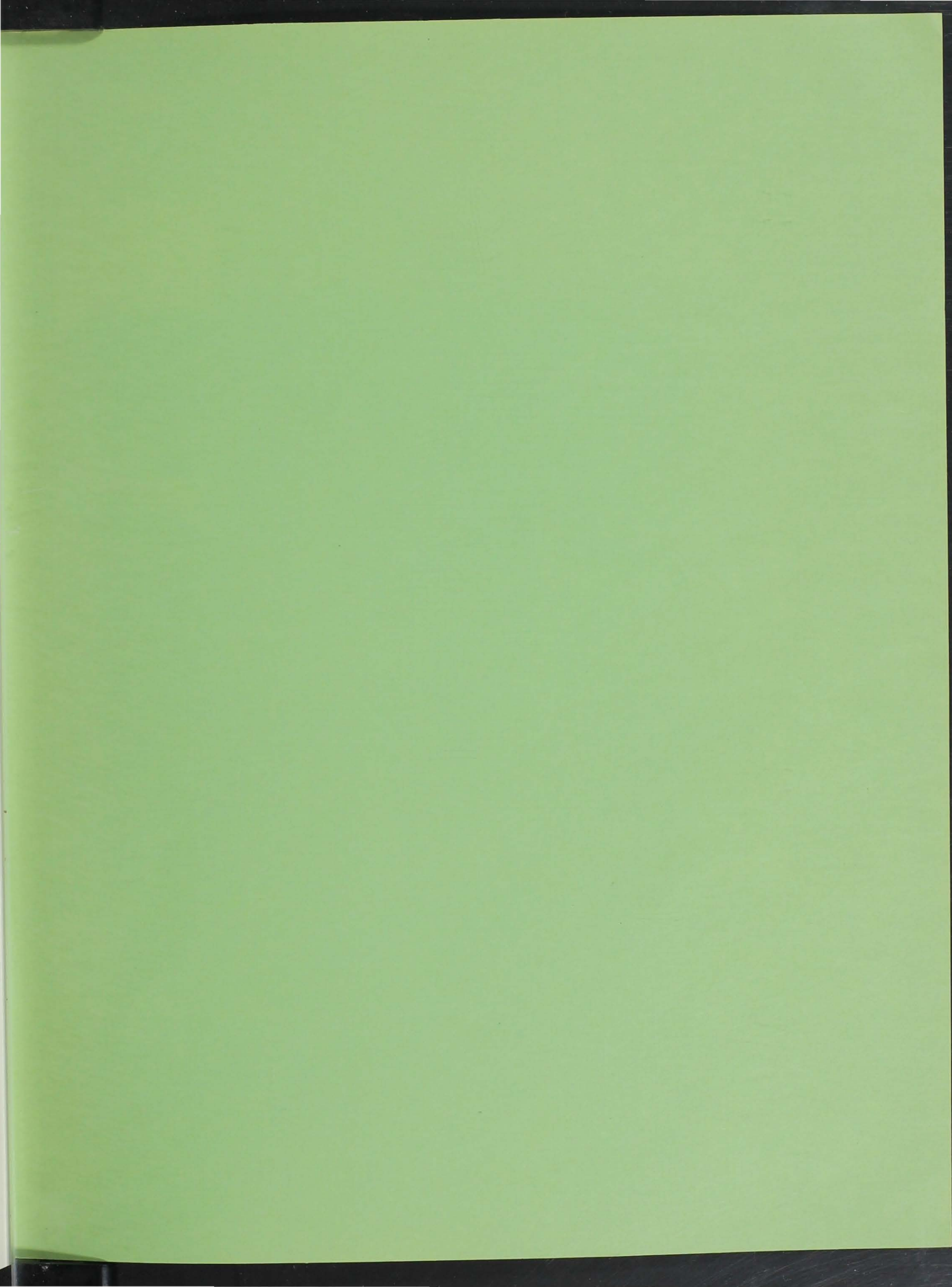
ed and taped) make good step down transformers for the operation of small motors, lights, heaters, etc.

Automobile wreckers and secondhand parts dealers are sources of some usable equipment. Automobile generators, when driven by a suitable motor through a V belt, make good sources of 6 v. or 12 v. direct current for electrolysis and similar purposes. Small motors from automobile heaters make usable stirrer motors. These can be fitted with inexpensive chucks obtainable from supply houses. Generator control relays can be rewound to make good relays for controlling constant temperature baths, etc. Junk storage batteries are usually a poor investment, as they are often too far gone to be of much value in a laboratory after having started a car for a year or more. Occasionally a good battery may be secured from a wrecked new model car. The electrolyte, which may have been spilled, can be easily replaced by a chemist.

Junked automatic washing machines are a prolific source of 115 v. motors as well as relays, solenoids, timers, gear boxes, circulating pumps, and thermostatic valves, all of which are useful for scientific purposes.

The control boxes for stokers and for some oil burners contain excellent power relays usually operating from small step down transformers. These are very useful for controlling heaters on ovens, constant temperature baths, and display equipment. Stoker controls also contain a very nice clock mechanism and contacts for turning current on and off on a time basis. The cams are usually set for operation at one-half-hour intervals, but additional cams may be added to shorten the time intervals.

The local telephone company may have surplus electrical equipment which they are willing to give to the school for physics and chemistry projects.



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